The Medical Epitome Series

PHYSICS
AND
INORGANIC CHEMISTRY

McGLANNAN

PEDERSEN

The Medical Epitome Series.

PHYSICS AND INORGANIC CHEMISTRY.

A MANUAL FOR STUDENTS AND PRACTITIONERS.

BY

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AUTHOR'S PREFACE.

In a book like this Manual of Inorganic Chemistry and Physics, new discoveries and novel theories can not be treated. The purpose is to set forth the accepted and proved facts forming the basis of the sciences in a manner which, in the author's opinion, will serve best for their clear and easy understanding by the student.

In preparing this volume I have consulted all available works on the subjects, drawing on them for the facts stated. In some places the phraseology of the masters has been copied when it has seemed that lucidity and terseness would be lost in any change. In this way I am under especial obligation to Professor William Simon, of Baltimore, who has permitted my drawing extensively on his "Manual of Chemistry" for inspiration, expressions, and illustrations.

To Dr. Pedersen, the Editor of this series, I owe much for valuable suggestion and supervision.

A. McG.

Baltimore, 1903.

EDITOR'S PREFACE.

In arranging for the editorship of *The Medical Epitome Series* the publishers established a few simple conditions, namely, that the Series as a whole should embrace the entire realm of medicine; that the individual volumes should authoritatively cover their respective subjects in all essentials; and that the maximum amount of information, in letter-press and engravings, should be given for a minimum price. It was the belief of publishers and editor alike that brief works of high character would render valuable service not only to students, but also to practitioners who might wish to refresh or supplement their knowledge to date.

To the authors the editor extends his heartiest thanks for their excellent work. They have fully justified his choice in inviting them to undertake a kind of literary task which is always difficult—namely, the combination of brevity, clearness, and comprehensiveness. The authors have shown a consistent interest in the work and an earnest endeavor to coöperate with the editor throughout the undertaking. Co-operation of this kind ought to result in useful books, in brief manuals as contradistinguished from mere compends. The editor desires at this opportunity to express his appre-

ciation of their helpfulness in the matter of producing the proper character of work.

In order to render the volumes suitable for quizzing, and yet preserve the continuity of the text unbroken by the interpolation of questions throughout the subject-matter, which has heretofore been the design in books of this type, all questions have been placed at the end of each chapter. This new arrangement, it is hoped, will be convenient alike to students and practitioners.

V. C. P.

NEW YORK, 1903.

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PHYSICS AND INORGANIC CHEMISTRY.

PART I.

PHYSICS.

CHAPTER I.

INTRODUCTION.

THE Sciences of Physics and Chemistry describe the different forms and conditions of matter, and the changes therein under various circumstances. When matter is altered in form without any modification in composition the process is *physical*, and studied by Physics. Changes involving alterations of the essential composition of matter are *chemical*, and treated of in Chemistry. The two series of changes are interdependent, and for correct understanding of either some knowledge of the other is necessary. Many chemical changes require for production the use of apparatus constructed on physical principles. Therefore a consideration of certain elements of physics will precede the study of chemistry.

DEFINITIONS.—Physics, that science which treats of matter and the forces and forms of energy of the universe, is divided into *general* and *applied physics*.

GENERAL PHYSICS treats of the phenomena of the action of force on matter, and the laws and the measure of the constants of these phenomena.

APPLIED PHYSICS is the correlation of the results of study in general physics with the data of special observation, for the purpose of explaining the phenomena of the universe.

The study of the constitution of matter and its properties, and of heat, light, and electricity, is included in general physics. Astronomy, geology, terrestrial magnetism, and meteorology are subdivisions of applied physics.

FUNDAMENTAL PROPERTIES OF MATTER.

DEFINITIONS.—Matter is anything that occupies space and is perceptible to our senses. All varieties of matter, even those which possess widely different characteristics, have in common certain properties, known as the *essential* or *fundamental properties of matter*. *Extension*, *Mobility*, *Divisibility*, *Elasticity*, *Gravitation*, *Porosity*, and *Indestructibility*, are the most important.

Extension is that property by which matter occupies space.

A BODY is a definite portion of matter—*e. g.*, a lump of coal, a pen, a brick. MASS is the quantity of matter possessed by a body. All bodies have length, breadth, and thickness.

A SUBSTANCE is a particular kind of matter—*e. g.*, silver, iron, and wood.

Mobility is that essential property of matter by which *position* may be changed, either that of one mass with reference to another mass, or that of particles to other particles within the mass.

Divisibility makes matter separable into smaller particles, as by cutting, crushing, tearing, and similar mechanical means, or by other agents, as heat. The extent of the subdivision depends upon the agent used—*e. g.*, ice may be broken with a hammer into several particles, one of which may be ground in a mortar into a fine powder; a particle of this placed under a microscope and heated will be converted into water, proving that it was capable of subdivision and consisted of smaller particles which under heat became more mobile. This drop of water heated in an empty flask will form steam, a gas, and completely fill it. These facts show that mechanical means cause only a certain degree of subdivision, and that the smallest perceptible particle consists of many much smaller particles. These smallest particles are called *molecules*.

Molecules are the smallest particles in which a substance may exist *as such*, or into which it may be *mechanically* or *physically* divided.

Atom is a term denoting the final subdivision of matter *chemically*—*i. e.*, *with a change in composition*. Molecules contain two or more atoms, but atoms are not further subdivisible. *Molecules* make up *substances*—*e. g.*, water is composed of molecules, each containing two atoms of hydrogen and one atom of oxygen (H_2O). Atoms are therefore the final unit of subdivision of the *elements*. A mass consists of an immense number of molecules separated from one another by intermolecular spaces.

MOLECULES and INTERMOLECULAR SPACES are proved to exist—in water, for example—by the fact that 1 c.c. of water boiled at ordinary atmospheric pressure forms 1800 c.c. of steam. If the 1 c.c. of water be converted into steam in a flask holding 1800 c.c., from which the air has been exhausted, and there be introduced into the flask a second and a third liquid, as alcohol and ether, the same quantity of these liquids evaporates each in the presence of the others as would have evaporated had the flask been empty and only one of the liquids put into it for

vaporization. The pressure in the flask rises with the introduction of each gas, because of the tension of each gas.

The GENERAL OCCURRENCE OF MOLECULES AND INTERMOLECULAR SPACES is shown by the above, and by the facts that carbon dioxide gas passes through red-hot iron, that water under pressure passes through gold, that no increase in volume of a fluid occurs when a salt is dissolved in it, etc. "A molecule is 0.00000006 millimeter in diameter. At first glance, the figure seems beyond human comprehension; yet these particles of matter are measurable, with shape, motion, and laws of action" (Sir W. Thompson).

The Law of Avogadro.

One of these laws is that of Avogadro: *Equal volumes of different gases at the same temperature and under the same pressure contain the same number of molecules.*

Gravitation is the attraction of masses for each other.

The Law of Newton.

Newton's law of gravitation is: *The force of attraction of bodies for each other is directly proportionate to their masses and inversely proportionate to the squares of their distance apart.*

A body thrown into the air falls to earth without exerting perceptible attraction for the earth because of the great difference in their relative masses. There is, however, a mutual attraction between the falling body and the earth, which may be shown by exact experiment. Other phenomena of gravitation are the flowing of water and the resistance of a body to being lifted or carried.

Weight is the quantity of mutual attraction between a body and the earth, and is expressed by a comparison or measurement of it with or by some standard weight, as pound, gram, etc.

Specific weight or **specific gravity** (*Sp. G.* or *S. G.*) is the ratio of the weight of a body compared with that of an equal volume of another substance, taken as a standard.

Density or **comparative mass** is the mass of a body compared with the mass of an equal volume of a standard body, and is often used for specific weight. *For solids and liquids the standard is water at the temperature of 15° C.; and for gases, either atmospheric air or hydrogen.*

I. To Determine the Specific Gravity of Solids.

1. SOLIDS HEAVIER THAN WATER.—First weigh it in air and then in water. The difference is the weight of the volume of water displaced. The weight in air is divided by this figure.

2. Solids Lighter than Water.—First weigh it in air, then in water with a sinker attached. The specific weight of the sinker is then determined and subtracted from that of both. The specific gravity is calculated as before.

3. Solids Soluble in Water.—The specific gravity is determined in a liquid having no solvent action, the specific gravity of which is known, and the weight of a corresponding volume of water is calculated.

II. To Determine the Specific Gravity of Liquids.

Weigh equal volumes of water and the given liquid and divide the latter by the former.

The **pyknometer** is a flask of thin glass holding accurately when completely filled a given volume at a definite temperature, and is much used for determining specific gravity of fluids.

The **hydrometer** is an instrument used for this purpose also, and is generally made of glass tubing, weighted below to maintain it upright when floating and with a stem above showing a scale. One form sinks in water to a mark on the scale—usually 1—and the other marks show at once the specific gravity of the fluid, according as the instrument floats above or below or at this level.

III. To Determine the Specific Gravity of Gases.

The fact that gases, when liberated, expand in all directions seems to indicate that their molecules are not attracted by the earth, and consequently without weight. This is not true. *A flask will weigh more filled with air than exhausted of air, and its weight will vary when successively containing different gases.*

The **atmosphere** is a layer of mixed gases, 50 to 100 miles thick, encircling the earth, and exerting pressure by its weight.

The **barometer** is an instrument for measuring atmospheric pressure.

A **mercury barometer,** the most common form, may be constructed by filling a glass tube, 1 meter long, closed at one end, with mercury, and inverting it in a vessel containing mercury. The mercury no longer fills the tube, but, leaving a vacuum above, falls to the height of about 30 inches, where it is held by the pressure of the atmosphere on the surface of the mercury in the vessel. Therefore the weight of the column of mercury 30 inches high must be equal to the weight of a column of air of the same diameter and as high as the layer of atmosphere is deep. The weight of a 30-inch column of mercury, 1 inch square, is 15 pounds. The atmospheric pressure is therefore 15 pounds to the square inch, which is borne without inconvenience because perfectly uniform in all directions. The atmospheric density and consequently the weight of mercury supported vary, chiefly because of atmospheric currents, temperature, and moisture. *The height of the mercury is lowered constantly with elevation above the sea-level.*

Porosity concerns the spaces between the molecules of a substance, called PORES, which may be large enough to be seen and felt, as in charcoal, sponge, etc. The proofs of the existence of molecules and intermolecular spaces, referred to above, are applicable to pores which cannot be seen, even under the microscope.

Surface and Surface-action.—The SURFACE is the perceptible limiting parts of a body and of its pores. The surfaces of certain substances show an attraction for certain others, known as SURFACE-ACTION or SURFACE-ATTRACTION. *Adhesion, capillary attraction, the absorbing power of solids and liquids, are phenomena depending on surface-action.*

Adhesion is the attraction of the surface of solids for some liquids, and the force joining the surfaces of bodies of different materials: wetting by water or alcohol, and gluing in the arts, are examples.

Capillary attraction is shown by the surface of a tube for the contained liquid. Whenever the liquid wets the tube, the fluid has a concave surface, that near the sides being drawn *up*. When the liquid does not wet the tube, as mercury, the surface is convex, that near the sides being drawn *down*. Capillary elevation or depression varies with the diameter of the tube, the temperature, and nature of the fluid. Narrower tubes and lower temperatures increase the elevation or depression. When capillary elevation occurs, the adhesion between the tube and liquid is greater than the cohesion of the liquid. The reverse exists with capillary depression.

Absorption.—Solids absorb gases, liquids, and solids held in solution by surface-action. A solid of known weight will show a decrease of weight while warm and return to its original weight on cooling. The loss of weight is due to the dissipation by heat of gases condensed on the surface.

When a solid, particularly if porous, like charcoal, is in contact with a mixture of liquids or a solution of several solids, the liquids or the dissolved solids are partly retained by it. The porous solid also retains more of one kind of liquid or solid than of another. The purification of water by charcoal-filtering, and the decolorization of sugar solution by bone-black, are due to this form of surface-attraction.

The surface-attraction of liquids for gases causes their condensation and absorption. The quantity of gas absorbed by a liquid varies for each gas and liquid; and is increased by low temperature and high pressure, and diminished by high temperature and low pressure.

Indestructibility is that property by which matter cannot be destroyed, although its form, appearance, and other properties may be changed in many ways. The burning of oil is simply a change in its form; gases are formed and diffused in the atmosphere, and thence absorbed by plants or otherwise changed, but never destroyed. *The total quantity of matter in the universe is constant.*

Elasticity causes matter to resist compression, traction, and torsion, and to return, in whole or part, to its former size when the force is removed. Rubber is a very elastic solid.

THE STATES OF MATTER.

Matter exists in three conditions or states of aggregation, viz., *solid*, *liquid*, and *gaseous*.

Convertibility of the States of Matter.

Certain substances may be changed from one state to another: solids liquefiable by heat are called *fusible*. Solids and liquids forming gases under heat and returning to solids and liquids, respectively, when cold, are termed *volatile*. Water is the only known substance existing in nature in all three states: solid as ice, fluid as water, and gaseous as atmospheric vapor under the sun-rays.

Solutions.—By the action of certain fluids, called *solvents*, solids and gases may be converted into liquids, producing *solutions*, and the solid or liquid is said to be *dissolved*. The solution may be physical or chemical. In PHYSICAL SOLUTIONS, after the removal of the solvent by evaporation or other means, the substance dissolved is recovered in its original condition. In CHEMICAL SOLUTIONS the dissolved substance in passing into solution is altered in composition. A SATURATED SOLUTION contains all of a given substance that can be taken up by the solvent at a given temperature.

The Solid State.

DEFINITION.—Solids are characterized by a *self-retaining figure*, due to a mutual attraction between the constituent particles, which gives definite size and shape, to alter which requires force. This attraction is called *cohesion*.

Crystals and Crystallization.

DEFINITIONS.—*Crystals* are solids, whose atoms or particles are arranged in a definite manner, and whose external appearance is a natural, regular figure, bounded by plane surfaces, related symmetrically by fixed laws.

Crystallization is the process of the formation of crystals. Its explanation assumes that particles attract one another in certain fixed directions.

Laws and Morphism of Crystallization.—This attraction can operate only under free motion of the particles, as after the solid has been rendered liquid (*melted*) or gaseous (*volatilized*) by heat or after solution in some appropriate agent (*dissolved*).

In the liquid or gaseous state the particles are subjected to their own attraction and reunite as crystals on removal of the cause of the change of state.

Many solids cannot form crystals; these substances are called *amorphous*, and are usually more complex in nature than those that crystallize. Some substances crystallize under certain conditions, but under others are amorphous. Certain substances form crystals of two or more shapes under different conditions of crystallization, and are called *dimorphous* or *polymorphous*. *The same substances under the same conditions always form the same type of crystal.* Two different substances forming the same variety of crystal are called *isomorphous*—*e. g.*, magnesium sulphate and zinc sulphate. Moreover, crystals of one such substance must develop in the solution of the other in order that both be called isomorphous. A crystal of zinc sulphate placed in a saturated solution of magnesium sulphate will be covered with crystals of magnesium sulphate. When a solution of isomorphous substances crystallizes, the crystals are made up of both substances.

When a solution containing two or more *non-isomorphous* (*heteromorphous*) substances, as sodium sulphate and sodium chloride, crystallizes, distinct crystals of each substance will be formed. The separation of solids by crystallization depends on this property.

Axes of Crystallization.—These are imaginary lines passing through the centre of the crystal and uniting *opposite, parallel sides* or *opposite angles*, assumed to exist, for defining the position of the planes and angles.

Water of crystallization is the water taken up in definite quantity by certain substances as a necessary element in the formation of crystals. *A given substance always requires the same quantity of water to develop the same form of crystal under the same conditions.*

Characteristic Properties of Solids.

Hardness, brittleness, tenacity, malleability, and ductility, are the most significant properties, and depend on differences in cohesion.

Hardness resists attempts to force a passage between or to displace the particles of a solid. A diamond is harder than glass.

Brittleness is the property of breaking easily under external force.

Tenacity resists attempts to pull the particles asunder. Steel wire is remarkably tenacious.

Malleability permits solids to be hammered or rolled into sheets. Gold is very malleable.

Ductility allows solids to be drawn into wire.

The Liquid State.

Liquids have *no self-subsistent figure*, consequently require a container, and present a horizontal surface. In liquids cohesion acts less energetically than in solids, permitting free motion of particles, but acts sufficiently to form drops.

The Gaseous State.

Gases have *no self-subsistent figure.* Their molecules move freely in every direction. Cohesion is entirely suspended. The molecules repel one another, and tend to spread in all directions. Gases require a closed vessel, on whose sides the molecular motion produces *pressure*, a characteristic of gases, known also as *tension* or *elasticity.*

The Law of Boyle or the Law of Mariotte.

The volume of a confined gas is inversely as the pressure, and the density and elasticity are directly as the pressure and inversely as the volume.

Diffusion is the gradual and spontaneous passage of one gas into another gas, one liquid into another, or of a dissolved substance into another liquid.

I. Diffusion of Gases.—If a bell-jar of oxygen be placed base to base over a bell-jar of nitrogen dioxide, the gases will pass into one another until a homogeneous mixture results, and this will occur even when they are separated by some porous substance. The rapidity of diffusion depends on the specific gravity of the gases, a specifically lighter gas diffusing more rapidly according to the law of Graham.

The Law of Graham.

The diffusion-velocity of a gas is inversely proportional to the square root of its density.

II. Diffusion of Liquids.—This may be observed by pouring water on glycerin. After standing a uniform mixture results. Water poured on a solution of salt will gradually mix with the latter into a uniform solution. If the two solutions are separated by a porous diaphragm, as a bladder or parchment-paper, the diffusion will occur as before, and the process is called *osmosis* or *dialysis.* The apparatus used is called a *dialyzer.*

Different substances behave variously during dialysis. In general, crystalline osmose more rapidly than amorphous substances. Those substances which do not dialyze are called *colloids*, and those which do, *crystalloids.*

III. Diffusion of Solids.—Perfect solution is first necessary in liquids which will diffuse when in contact. The diffusion of the solution then occurs according to the laws for this process among liquids.

FORCE AND ENERGY.

Force is the action of matter on matter, changes or arrests motion, and is a manifestation of energy, originated in various ways.

Energy, a property of all matter, is the power of performing work—*i. e.*, of transferring motion from one body to another.

Forms and Convertibility of Energy.

Energy exists in different forms: the motion of bodies is *external* with reference to other bodies, or *internal* with reference to their own atoms. Heat, light, electricity, and chemical action, are all forms of energy. Energy is *potential* or *kinetic*.

Potential energy is that of position, or restrained motion, capable of being converted entirely into *active* or *kinetic energy*. The energy of water is potential in a reservoir and kinetic in turning a mill-wheel when released.

The Law of the Conservation of Energy.

The QUANTITY OF ENERGY in the universe is constant, although its varieties are infinite. The total potential energy of the reservoir will become the same kinetic total in the mill-machinery and the heat of friction. This total kinetic energy is reconverted into an equal potential-total in doing the work, and stored up somewhere for future change. The principal stores of energy available for human use are the combustibles, moving water or water at an elevation, atmospheric currents, muscles of animals, the tides, and direct solar radiation. *The source of all these forms of energy, except that of the tides, is the radiant energy of the sun.*

THE INDESTRUCTIBILITY OF ENERGY.—One form of energy may be converted into another without loss. Heat may be converted into motion, motion into chemical energy or electrical energy, electrical energy into heat. This fact is spoken of as the law of the conservation of energy.

QUESTIONS.

What is matter?
Define five fundamental properties of matter.
What are the characteristics of the three states of aggregation?
Define force and energy.
What is meant by potential energy and kinetic energy?
What is meant by a fusible substance?
What is meant by a volatile substance?
What are the differences between amorphous, polymorphous, and isomorphous substances?
What is water of crystallization?
What are the axes of crystals?
Define the characteristic properties of solids?
Define cohesion, adhesion, and gravitation.
State the law of Boyle or Mariotte.
What is a mass, and what is a molecule?
Explain the existence of molecules and intermolecular spaces.
State the law of Avogadro.
Define weight and specific weight.
How may a specific weight be determined of a liquid, a gas, a solid heavier than water, a solid lighter than water, and a solid soluble in water?

What is an hydrometer?
What is a pyknometer?
What is the ordinary atmospheric pressure?
How is atmospheric pressure determined?
What are surface-action and capillary attraction?
What is diffusion?
What is dialysis?
What is the difference between colloids and crystalloids?
State the law of Graham.
What is the law of the conservation of energy?
What are meant by solution and by saturated solution?

CHAPTER II.

HEAT.

THE THEORY OF MOTION IN HEAT.

THE molecules of all substances are in constant motion: those of solids being held in a fixed position can vibrate only forward and backward, those of liquids are less restricted, and those of gases are free in all directions. All molecules thus have *kinetic* energy, and those of liquids and solids have also *potential* energy, due to their molecular cohesion, to overcome which requires energy from outside sources. When this potential becomes kinetic energy and the molecules vibrate with a certain rapidity, that form of kinetic energy called heat is developed.

THE SOURCES OF HEAT.

The sun is the principal and the molten interior of the earth, friction, compression, electric currents, chemical action, as ordinary combustion, and animal life, are other sources of heat.

THE GENERATION OF HEAT.

The common method of generating heat is by combustion (see the chapter in Chemistry on this subject).

The electric furnace produces temperatures much higher than formerly possible; sufficiently high to liquefy or volatilize all substances.

THE EFFECTS OF HEAT.

Heat-energy changes the position of the molecules of a body, or increases their motion, or both. With this increase of motion—*i. e.*, kinetic energy—a rise in *temperature* occurs measurable by thermometers.

Temperature is the ability of a body to give or receive heat.

The effect of heat on molecular motion and position is illustrated by its action on ice. A thermometer in a vessel containing melting ice records 0° C. (32° F.). Heat applied to the vessel will melt the ice by overcoming the potential energy of its molecular cohesion. Until the last particle of ice has disappeared the same temperature is recorded. After that, under continued heat the temperature will rise to 100° C. (212° F.), when the water will be converted into gas (steam), because the potential energy of the cohesion has been entirely overcome, and as steam the molecules move freely in all directions. Here the thermometer's record remains stationary until the water has disappeared.

Latent Heat.

PRINCIPLE.—*Heat-energy was used in converting the ice into water, and the water into steam, and yet is not recorded by the thermometer.*

DEFINITION.—This *unrecorded* heat—the energy required to overcome the potential energy of molecular cohesion of solids and liquids—is called *latent heat.*

Sensible or Free Heat.

LATENT heat becomes *sensible* or *free heat* when the process is reversed, as in heating by steam condensation.

Expansion and Contraction.—As a rule, bodies expand when heated. A bar of metal which gauges accurately when cold is too large when heated; and on cooling again gauges. A glass bulb with a narrow neck filled with liquid, shows expansion of the liquid up the neck under heat and a contraction to its original volume on cooling.

If an open flask of air be inverted under water and heated, the air will expand, and finally will bubble out through the water. On cooling, the water will rise into the flask, replacing the escaped air.

Exceptions to Expansion.—Notable exceptions to this rule are: water contracts when heated from 0° C. (32° F.) to 3.9° C. (39° F.), but expands when further heated; and cast-iron, sulphur, and type-metal, all expand at the point of solidification.

The *degree of expansion* varies with the different states of matter, being least in solids and greatest in gases. Metals expand more than most other solids; solids and liquids vary with each particular substance, but gases expand with fixed regularity, which was determined by Charles.

The Law of Charles.

Under constant pressure the volume of a gas increases or decreases directly as the temperature. With a constant volume the pressure increases or decreases directly as the temperature.

The expansion or contraction of gases is 0.3665 per cent. or $\frac{1}{273}$ of their volume at 0° C. for each degree Centigrade.

Expansion and contraction are respectively due to increase or diminution in rapidity of molecular motion.

The limit of increased motion is unknown, but that of decreased motion is probably determined. According to the law of Charles, the volume of a gas at 0° C. will increase $\frac{1}{273}$ for each degree centigrade raised. Accordingly a mass of air at 0° C. heated to 273° C. will double its volume. If the temperature be reduced, at —273° C. the air would mathematically have no volume. What actually would occur in reaching this low temperature is, the air would be liquefied, then solidified, and finally at —273° C. become a compact mass at absolute molecular rest.

Absolute Zero and Absolute Temperature.

DEFINITION.—This point of no motion or no heat is called *absolute zero*, and temperature reckoned from it, *absolute temperature.*

The absolute temperature may be found by adding 273 to the reading on a Centigrade and 491 to that on a Fahrenheit thermometer. The fractions $\frac{1}{273}$ *and* $\frac{1}{491}$ *are known as the coefficients of expansion for Centigrade and Fahrenheit degrees, respectively.*

It is often necessary to reduce the volume of a gas at any temperature and pressure to the volume it would occupy at 0° C. and 760 mm. barometric pressure—*i. e.*, at the *standard temperature* and *pressure.*

1. To Reduce the Volume of a Gas to that at 0° C.

The volume of a gas is proportional to its absolute temperature. Therefore to reduce the volume of any gas to 0° C., find its absolute temperature and calculate by simple proportion—*e. g.*, if a gas measures 119 c.c. at 36.4° C., its volume at 0° C. is found from the proportion:

$$119 : (36.4^\circ + 273^\circ) :: x : (0^\circ + 273^\circ)$$

$$119 : 309.4 :: x : 273$$

$$x = \frac{119 \times 273}{309.4} \text{ or } x = 105 \text{ c.c.}$$

In this reduction the pressure is supposed to remain constant.

2. To Reduce the Pressure of a Gas to Normal.

Use BOYLE'S LAW—viz., The product of the pressure multiplied by the corresponding volume of a gas is always constant when the temperature is the same. Expressed in equation, Boyle's law is, $PV = pv$.

If in the above case the 105 c.c. were under 755 mm. pressure, the volume at 760 mm. is found by the equation:

$$105 \times 755 = x \times 760$$

$$x = \frac{105 \times 755}{760} = 101.67 \text{ c.c.}$$

Evaporation.—*Many liquids and solids are converted into gases at nearly all temperatures*—*e. g.*, water and camphor vaporize at ordinary temperatures. The fact is explained by assuming that while in rapid motion some molecules get beyond the attraction of the surrounding molecules and are mixed with the atmospheric molecules.

Boiling.—This evaporation must not be confounded with boiling, which is the rapid conversion of a liquid into a gas, at a fixed temperature, with ebullition, which is the phenomena produced by the formation of gas in the liquid. This fixed temperature is the *boiling-point* of the liquid, and is the highest point to which it may be heated under normal atmospheric pressure.

Pressure materially affects the boiling-point. Water boils at 100° C. or 212° F. under 1 atmosphere. Increased pressure, as in a steam boiler, raises the boiling-point—pressure of 2 atmospheres raises it to 122° C. (251° F.). Reduced pressure, as on a mountain or in a vacuum, lowers the boiling-point.

PAPIN'S DIGESTER is a closed boiler for heating water under pressure and raising its boiling-point. At the high temperature under 4 or 5 atmospheres albuminoid substances, ordinarily insoluble, are quickly dissolved.

Liquefaction and Vaporization.—Solids are liquefied or vaporized by heat. Most solids form liquids and then gases at high temperatures; while others—*e. g.*, calomel—pass directly into a gas.

The temperature at which a solid is liquefied is called the POINT OF LIQUEFACTION (melting-point), and that at which it is vaporized, the POINT OF VOLATILIZATION.

Condensation.—The *condensation-point* is the temperature at which a gas becomes liquid. For the liquefaction of substances which are gases at all ordinary temperatures high pressure and low temperature are combined.

Critical Temperature.

For each gas there is a limit of temperature, called the *critical temperature*, above which it cannot be liquefied by any pressure however great. The failure of former attempts to liquefy many gases was due to the fact that although the pressure used was more than sufficient, the critical temperature was not reached.

Solidification.—FREEZING-POINT is the temperature at which a liquid becomes solid.

When a solid is dissolved in a liquid the boiling- and freezing-points will vary according to the concentration of the solution.

The addition of salt to water for culinary purposes, in high altitudes, is to provide a sufficiently high temperature to cook the food. The melting of ice by salt depends on the much lower freezing-point of a solution of salt than of pure water. The fall in freezing-point is nearly constant for each molecule of solid dissolved in 100 molecules of liquid.

Cryoscopy.—By this process the relation of simple to complex molecules in the urine is determined by the freezing-points of urine and of a solution of sodium chloride.

Distillation is the conversion of a liquid into a gas and the reconversion of the gas into a liquid.

Sublimation is the conversion of a solid into a gas and the reconversion of the gas into a solid.

Destructive distillation is the heating of non-volatile organic matter in such a manner that air is excluded and the organic matter is decomposed.

Fractional distillation is a distillation in which the receivers are changed with increasing temperatures of the liquid to be distilled. By this process liquids of different boiling-points may be separated from a mixture.

THE OBSERVATION OF HEAT.

The Thermometer.

DEFINITION.—These are instruments for measuring variations in *heat-intensity.*

PRINCIPLE.—The construction of thermometers depends upon expansion. The commonest are the *mercury* thermometers.

CONSTRUCTION.—Mercury is introduced into a capillary glass tube with a bulb, and boiled. The tube is then sealed. The instrument is then placed in boiling water and the boiling-point of water marked, next in melting ice and the freezing-point of water noted. On a *Centigrade* thermometer the distance between freezing-point and boiling-point is divided into 100 degrees; on a *Fahrenheit* thermometer, into 180 degrees: therefore, 1 degree C. = 1.8 degrees F. Fahrenheit did not begin to count with the freezing-point of water, but 32 degrees below this point; consequently on the Fahrenheit scale the freezing-point of water is 32° and the boiling-point 212°; on the Centigrade these points are 0° and 100°, respectively.

Degrees of heat below 0° in either scale are indicated by the *minus* sign—*e. g.*, —40° F., —20° C.

To convert Centigrade degrees into Fahrenheit, multiply by 1.8 *and add* 32. *To convert Fahrenheit into Centigrade, subtract* 32 *and divide by* 1.8.

Specific Heat.

This is the *heat-quantity* required to raise a definite weight of a substance a definite number of degrees compared with that required to raise the same weight of water the same number of degrees.

The Calorimeter.

DEFINITION.—This instrument measures the quantity of heat generated by the combustion of a substance.

PRINCIPLE.—The instruments are constructed to burn a definite quantity of a substance in oxygen under pressure, in a chamber surrounded by a known quantity of water, the rise of whose temperature is the basis for calculation.

The **calorie** is the unit of measure for heat-quantity.

The LARGE CALORIE is *the quantity of heat required to raise* 1 *kilogram of water* 1 degree *C., or about* 1 *pound of water* 4 degrees *F.*

The SMALL CALORIE *raises* 1 *gram of water* 1 degree *C.*

Joule's Equivalent of Heat.

Heat and work being different forms of energy, are interconvertible, 1 calorie of heat is equivalent to the *work-energy* of raising 426.5 kilograms 1 meter high or 3077.6 pounds 1 foot high. *Joule's equivalent is, therefore,* 426.5 *kilogram-meters, or* 3077.6 *foot-pounds.*

THE TRANSFERRENCE OF HEAT.

Heat may be transferred from one substance to another by *conduction, convection,* or *radiation.*

Conduction is transferrence from one molecule directly into its neighbor.

Convection is the transfer by ascending and descending currents, equalizing the temperature. If a piece of ice be fixed in the bottom of a test-tube, filled with water, the upper portion of the liquid may be boiled without melting the ice. If, however, the heat be applied to the bottom of the tube, currents are set up in the liquid, the temperature of which remains at 0° C. until the ice has melted and then gradually rises. *The atmosphere is heated chiefly by convection.* The air in contract with the earth is warmed, and rises, creating a current. The heated air in rising also moves laterally, and in this way produces winds.

Radiation.—To explain the transfer of heat by *radiation*, the existence of the hypothetical agent ether is assumed, a medium of extreme tenuity and elasticity, supposed to pervade the universe and all matter. The rapid molecular vibrations of a heated body set up vibrations in it, similar to those in a pool of water by a stone thrown into it. The length and speed of these vibrations of ether may be measured accurately.

When the heat waves of ether strike a body, they may be absorbed, reflected, transmitted, or diffused. When *absorbed*, they cause more rapid molecular movement of the absorbing body, and therefore increase its heat.

Transfer of heat by ether waves is heating by radiation, and is the method by which the sun's heat reaches the earth.

QUESTIONS.

What are the principal sources of heat?
What is the effect of heat on a substance?
What are latent heat, free heat, and specific heat?
State the law of Charles.
What is meant by absolute zero and absolute temperature?
What is the absolute temperature corresponding to —12° C., to 33° C., to 55° F., to —25° C.?
What are the normal temperature and pressure for measuring gas-volume?
What will be the volume of a gas at normal temperature and pressure which measures 66 c.c. at 54.6° C. under a pressure of 758 mm.?
What is meant by boiling-point, melting-point, and freezing-point?
What is the influence of pressure on boiling-point?
What is Papin's digester?
What is meant by cryoscopy, and on what principle does it depend?
What is meant by distillation, sublimation, and destructive distillation?
Explain the construction of a mercury thermometer.
How many degrees Fahrenheit correspond to 45° C., to 10° C., to —20° C., to —40° C.?
How many degrees Centigrade correspond to 50° F., to 104° F., to —23° F., to —40° F.?
What is a calorie, and what is the relation of 1 calorie of heat to work?
What is meant by conduction, convection, and radiation of heat?
What is the critical temperature of a gas?
What is the value of the large calorie? Of the small calorie?
State Joule's equivalent of heat.

CHAPTER III.

LIGHT.

DEFINITION.—Light is the visible form of energy.

THE THEORY OF MOTION IN LIGHT.

Like heat, light is produced by increased molecular motion, and transmitted by vibrations in the ether. An iron bar slightly heated emits heat, and highly heated, both heat and light. Greater heat causes increased molecular movement, consequently greater ethereal vibration, thus producing light.

The Relation of Thermal, Photic, and Chemical Action.

That light is different from heat may be shown by the differences of action in producing chemical changes: *Photographs* result from light acting on silver salts.

Therefore molecular motion sets up vibrations in ether, and produces

at the same time heat, light, or chemical action. All are different manifestations of radiant energy.

The Character of Light Waves.

VELOCITY.—All these ether waves travel through space at the velocity of 300,000 kilometers (186,000 miles) per second, but those producing such various results differ from one another in *length* and *height* (*frequency* and *amplitude of vibration*).

VISIBILITY.—Waves measuring in length from 650 to 442 millionths of a millimeter, and making between 477,000,000 to 699,000,000 million vibrations per second, produce the sensation of light.

COLOR.—According to their frequency, these waves produce sensations of color; the waves of least frequency and greatest length produce the sensation of red. With increase of frequency, sensations of orange, yellow, green, blue, indigo, and violet light appear successively.

Infra-red waves are those of too limited frequency to be visible, and have great heating power.

Actinic or *ultra-violet rays* are the invisible rays of greater frequency than violet vibrations, and have high chemical action.

THE ABSORPTION OF LIGHT.

Day-light (*white light*) is a mixture of waves of different length and height, affecting the eye simultaneously, so that no color predominates.

SPECIFIC COLOR.—Objects show a specific color under white light, not through any inherent color, but because the light may be entirely or partially *absorbed* or *reflected*. Absorption of all the rays gives *black*. Of all except the violet, makes violet. The color of objects is an effect of light, without it all objects are of no color—*i. e.*, black. Light itself is invisible. Consciousness of it is not by *seeing* its essence, but some object emitting or intercepting it.

Transparent bodies *readily* and *completely* transmit light, *translucent* bodies only *partially*, and *opaque* bodies *not at all*.

Fluorescence is the property of certain transparent substances of becoming self-luminous while exposed to direct light.

Phosphorescence is the property of luminosity without combustion and without exposure to light.

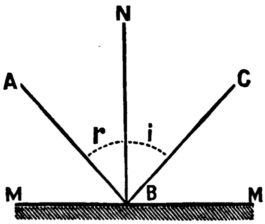
THE REFLECTION OF LIGHT.

Light always travels through a homogeneous medium in a straight line. A narrow cylinder of light is called a *ray*, *pencil*, or *beam*. A ray falling on a highly polished surface (mirror) is intercepted, and *reflected* (glances off).

The Law of Reflection.

With plane mirrors the angle of incidence is always equal to the angle of reflection (Fig. 1). The angle *i* between the incident ray *CB* and the perpendicular *NB* is equal to the angle *r* between the reflected ray *AB* and the perpendicular. *Heat is reflected according to the same law.*

FIG. 1.



Reflection.

Images in mirrors depend on reflection. A perfectly smooth plane reflecting surface is invisible, reflected images only are seen in it. Most objects are bounded by more or less rough surfaces, whose irregularities reflect, diffuse, or scatter light, and render the object visible in all directions.

THE REFRACTION OF LIGHT.

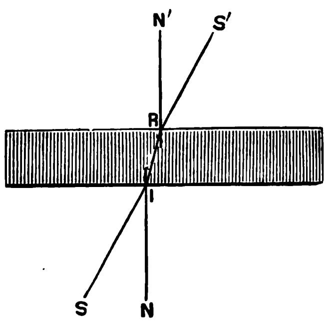
DEFINITION.—Refraction is the change of direction which a ray of light undergoes when passing from a medium of one density to another of different density—*e. g.*, air and glass. If the ray strike *parallel surfaces perpendicularly*, there is no refraction. If the ray does not enter *perpendicularly*, it bends out of its course and moves in a straight line in the second medium in a direction different from that in the first.

A straight stick held obliquely in clear water appears bent at the point of immersion, because of refraction.

The Law of Refraction.

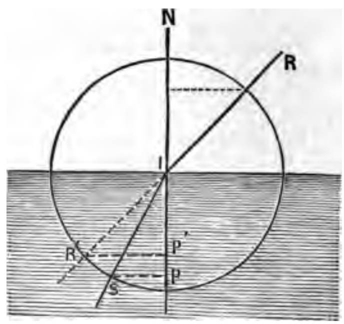
The ray *SI* entering a denser medium (a glass plate with parallel sides) bends *toward* the perpendicular *N′R*. On re-entering the air

FIG. 2.



Refraction through parallel plate.

FIG. 3.



Law of refraction.

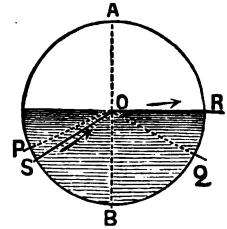
it again bends in the direction RS', *away* from the perpendicular, *to such an extent that extensions of SI and RS' are parallel* (Fig. 2.)

The ray RI passing from air into water bends and takes the direction SI. The angle RIN is the *angle of incidence*, and SIN the *angle of refraction* (Fig. 3).

The ratio between the sines of the angles of incidence and refraction is constant for the same media.

When one of these media is air, this ratio is called the **index of refraction** of the other medium—*e. g.*, if a ray of light pass into water, the ratio of the sines is 4 : 3. Therefore the *index of refraction of water* is $\frac{4}{3} = 1.33 +$.

FIG. 4.



Critical angle.

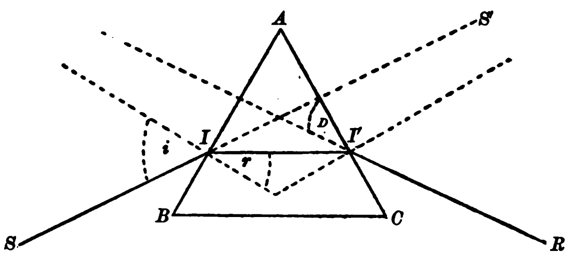
The Critical Angle.

As a ray emerging from denser into rarer media is refracted *from* the perpendicular, there is a limit of size of the incident angle permitting the refracted ray to emerge, called the *critical angle.* SOB is the critical angle, the largest allowing the ray SO to emerge into the air. The ray PO is reflected back into the denser medium as OQ (Fig. 4).

The Refraction of Light through Prisms.

DEFINITION.—In optics a *prism* is any transparent medium comprised between two planes inclined to each other, their intersection is the *edge*, and the angle between them, *refracting angle.* Triangular glass prisms are generally used (Fig. 5). ABC is the section of a

FIG. 5.



Refraction through a prism.

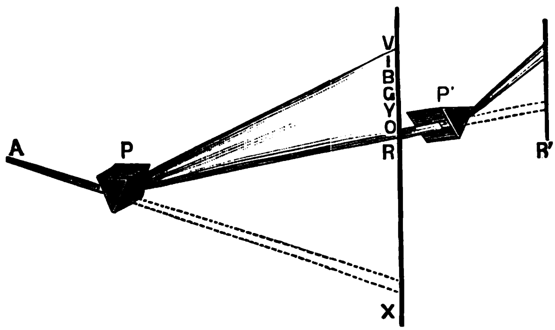
prism, A the apex, and BC the base. A ray SI falling on the prism will not pass through in a straight line SIS', but bends twice, in

accordance with laws of refraction, first from I to I' and then to R, in degree according to (1) the prism-angle, (2) its refractive index, and (3) the angle of incidence. In the figure i is the angle of incidence, r that of refraction, and D is the angle of deviation.

The Dispersion of Light through Prisms.

DEFINITION.—Dispersion is the spreading out and separation of light after its refraction into its constituent colors. In Fig. 5 the ray is represented by a line throughout. In reality the beam does not come through a prism as white light, but its constituent colors are refracted at different angles, giving rise to a *band of light* containing all the colors of the rainbow (violet, indigo, blue, green, yellow, orange, and red). Violet rays are refracted most, and red least. These colors (called the PRIMITIVE COLORS) gradually merge into each other, forming the

FIG. 6.



Prismatic spectrum.

prismatic spectrum. In Fig. 6 A represents a ray which unbent would strike the screen at X, but in passing through the prism P is refracted, dispersed, and forms the spectrum. The colors are indicated by their initial letters. If the ray on emerging is passed through a second prism, P', greater dispersion will occur.

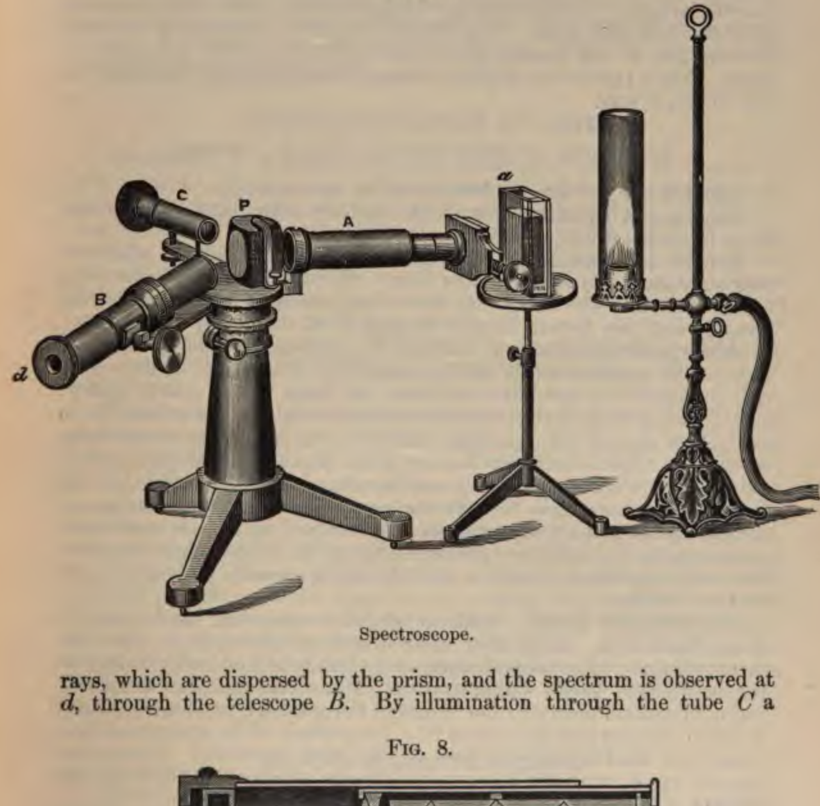
Achromatic prisms are constructed to refract but not decompose light, and consist of substances having different indexes of refraction and dispersion, combined to neutralize each other—*e. g.*, *crown* glass has its dispersion neutralized by that of *flint* glass.

The Spectroscope.

This instrument is for observing the spectrum. Fig. 7 represents the single-prism type. Its parts are the prism P, the tube A, at whose dis-

tal end light is admited through a slit controlled by a screen. A convex lens, the *collimator*, at the other end, collects the light in nearly parallel

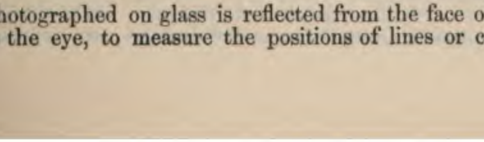
FIG. 7.



Spectroscope.

rays, which are dispersed by the prism, and the spectrum is observed at *d*, through the telescope *B*. By illumination through the tube *C* a

FIG. 8.



Direct-vision spectroscope.

fine-scale photographed on glass is reflected from the face of the prism down *B*, to the eye, to measure the positions of lines or colors. The

glass vessel *a*, having parallel sides, contains liquids for spectroscopy. The direct-vision spectroscope (Fig. 8) is more easily portable, consists of a cylindrical tube containing an ocular lens and from 3 to 7 prisms in a draw-tube, an adjustable slit, and a collimator lens between the slit and the prisms. The prisms placed opposite one another are alternately of *crown* and *flint* glass. The dispersive power of the latter is nearly double that of the former, while their refractive power is nearly the same: thus a ray is very slightly deviated, but sufficiently dispersed for spectroscopic uses.

THE TYPES OF SPECTROSCOPIC IMAGE (SPECTRA).

Spectra are the images observed in the spectroscope.

Continuous spectra show all the primary colors, and are observed from luminous solids or liquids.

Bright-line spectra are observed from luminous gases, as when certain substances are volatilized by heat. *Bright lines in different colors of the spectrum (the remaining colors being invisible) vary for each element, so that any element may be detected by its own line-spectrum.*

Absorption-spectra are obtained by interrupting the passage of light to the spectroscope by certain media. If a crystal of sodium chloride be heated in a non-luminous flame, the flame will be colored yellow, and if such a flame be examined spectroscopically a bright-yellow line in the yellow part of the spectrum will be seen (the other colors being invisible). *This is the bright-line spectrum of sodium.* If the spectroscope obtain a continuous spectrum from an electric light, and then the sodium flame be brought between the electric light and the spectroscope, the previously continuous spectrum is interrupted by a black band (*occupying exactly the position of the yellow line of the sodium spectrum*, called the *absorption-band of sodium*, and the spectrum is the *absorption-spectrum of sodium.*

Frauenhofer Lines.—Sunlight shows a spectrum interrupted by *many* black lines, called after their discoverer *Frauenhofer lines*, the more prominent ones being designated by letters as landmarks in describing positions of bands within spectra. These lines are due to volatilization of certain metals in the sun, and the presence or absence of any metal in the sun may be proved by a comparison of its absorption-spectrum with the Frauenhofer lines of the solar spectrum. Absorption-spectra are of great assistance in recognizing solutions of alkaloids, like quinine, and of blood and blood-coloring matter in solution.

THE DOUBLE REFRACTION OF LIGHT.

This is the power of refracting light in two directions. Printing viewed through glass is not altered, but through certain crystals (*calcium carbonate—i. e., Iceland spar*) is duplicated. Light through a pinhole in a card observed through a crystal of Iceland spar, appears as through

each of *two* holes. If the crystal be rotated in a plane parallel to the card, one hole appears to rotate around the other, which remains nearly stationary. From these facts it is evident that the entering ray is split into two, one, called the *ordinary*, obeying the laws of refraction already mentioned, while the other, disregarding them, is called the *extraordinary ray*. Thus are the phenomena of *double refraction* constituted. The OPTIC AXIS OF SUCH A CRYSTAL is *the line through which double refraction is suspended.*

THE POLARIZATION OF LIGHT.

Tourmaline is a double-refracting substance, adding the power of *absorbing the ordinary ray*. If two tourmaline plates cut parallel to the axis of the crystal are joined in a crossed position, light through the first is cut off by the second. If the two plates be turned on one another, they are most transparent in *two* positions, *one*, *the normal relative position of the plates in the crystal, and the other* 180 *degrees from this*. The combination is entirely dark in *two* positions, at right angles to the original position in the crystals, and intermediate positions show partial darkness. This behavior shows that light after passing through tourmaline exists in a peculiar condition. It is PLANE-POLARIZED.

DEFINITION.—Polarized light undulates only in the one plane, while ordinary light undulates in all planes parallel to the line of propagation. The ether undulations may be compared with those of a cord fastened at *one end* and shaken at the other. It will then undulate in a horizontal, oblique, or vertical plane, according to the direction in which the hand moves. A grating of *vertical* rods will permit the passage of all *vertical* planes, but intercepts all others. If the rods of two gratings are in the same plane, any vibration in a plane parallel to the rods of one will pass through the other—*e. g.*, a string may be made to vibrate through both gratings *parallel* to the rods. When the rods of the gratings are at right angles this becomes impossible. A tourmaline plate will thus transmit undulations in a plane *parallel to its axis*, but absorbs those which are in planes at *right angles to its axis*. The rays passing through produce the polarized light.

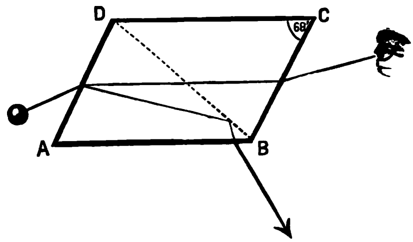
The Polariscopic Instruments.

The **polariscope** studies the effects of polarized light. The **polarimeter** measures degrees of polarization. The **saccharimeter** is a special polarimeter for the *quantitative* estimation of sugar.

CONSTRUCTION.—The *polariscope* consists of an opaque tube, with eye-piece and two polarizing crystals (the *polarizer* and the *analyzer*). In the *polarimeter* the analyzer is fixed in a rotating tube with an indicator and scale to show the degree of rotation. The analyzer and polarizer generally used are *Nicol's prisms*, constructed by sawing through a

crystal of Iceland spar, from one obtuse angle to another, polishing the cut surfaces and cementing with Canada balsam. The sides of the crystal are coated with lampblack. The ray entering at the side AD (Fig. 9) is doubly refracted. The *ordinary* ray reaches the balsam at the critical angle, is reflected back to the side AB, and absorbed by the lampblack. The *extraordinary* ray passing into a second Nicol s prism *parallel* to the first (as though a continuance of it) passes through *unchanged.* If the second prism be at right angles to the first—*i. e.*, *the prisms crossed*—the ray becomes in the second prism an *ordinary* one, is reflected out by the balsam and absorbed by the lampblack, and *no* light passes through. In *intermediate positions* the *extraordinary* ray is decomposed in the second prism into ordinary and extraordinary rays; the *former* is reflected out and the *latter* passes through. The amount of light increases as the two crystals approach parallelism.

FIG. 9.



Nicol's prism.

Thin sections of crystals of different crystallographic systems show peculiar rings or bands of colors in the polariscope. Advantage is taken of this fact in examining minerals and salts, particularly precious stones. Solutions of certain organic substances have the peculiar property of deviating the plane of polarization—that is, of turning the direction of the vibration of the light. If the prisms of a polariscope be crossed so that no light comes through, and a solution of one of the substances (for example, sugar) be placed between the analyzer and the polarizer, it will be observed that some light now passes. The sugar turns the plane of the light from the polarizer, and the effect is the same as though the analyzer had been rotated. OPTICALLY ACTIVE SUBSTANCES have this effect, and according as the *analyzer* must be turned to the right or left in order to restore darkness are DEXTRO- or LÆVO-ROTATORY. The degree of rotation of the plane of polarization is measured by the extent to which the second prism must be rotated to obtain darkness.

Specific rotation *of a substance is the degree of rotation required to restore darkness when a tube* 1 *decimeter long, filled with a solution of* 1 *gram of the substance in* 1 *cubic centimeter of solution is introduced between the prisms.* In most polarizers a crystal of quartz or some similar material is introduced to produce sharply defined color-tints, or unequal illumination of the halves of the field, or in some other way to provide points of comparison.

The Refraction of Light through Lenses.

DEFINITION.—A lens is a transparent medium (usually glass), bounded by two curved or one curved and one plane surface, and used to cause systematic *convergence* or *divergence* of light passing through it.

The Forms of Lens.

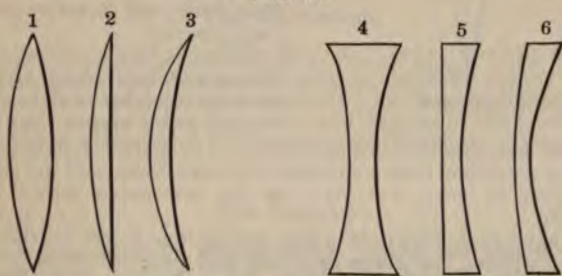
Lenses are of two types: (1) **convex, magnifying,** or **converging** lenses, thicker at the centre than edges; and (2) **concave, diverging,** or **reducing** lenses, thickest at the edge.

Each type has three varieties (Fig. 10).

(1) **Double convex** or **biconvex**, bounded by two curved faces whose radii are opposed and overlapping.

(2) **Planoconvex,** bounded by a plane and a spherical surface.

FIG. 10.



Forms of convex and concave lenses.

(3) **Meniscus,** bounded by curved surfaces whose radii are opposed but overlap.

(4) **Double concave** or **biconcave,** having both surfaces concave.

(5) **Planoconcave,** having one plane and one concave surface.

(6) **Concavoconvex,** one surface concave and the other convex. Here the radii though opposed do not overlap.

The AXIS OF THE LENS is the line piercing the centres of curvature of the two curved surfaces, or is the perpendicular to the plane face from the centre of curvature of the spherical face, when one surface is plane.

The DIAMETER passes from one point of the circumference to that exactly opposite, perpendicularly to the axis.

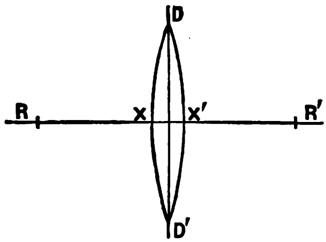
The OPTICAL CENTRE is a point on the axis so taken that every line drawn through it pierces parallel elements of the two surfaces. In Fig. 11 the line Rx', describing the curve $Rx'D'$, is the RADIUS OF CURVATURE; the point R, about which this revolves, the CENTRE OF CURVA-

TURE; the line xx', the PRINCIPAL AXIS OF THE LENS; and DD', the diameter.

The PRINCIPAL FOCUS is the point at which a convex lens converges rays parallel to its axis. The distance from the principal focus to the optical centre is the same on both sides of the lens, and depends on the material of the lens and the radii of the curved surfaces.

THE REAL FOCUS.—Rays from a source *beyond* the principal focus are collected to a *real focus*, *beyond* the principal focus on the *other* side.

FIG. 11.



Convex lens.

THE VIRTUAL FOCUS.—Rays from a source *at* the principal focus emerge *parallel*; but from one between the principal focus and the lens, converge, not to a *real* but a *virtual* focus, as from a point *behind* the source of light.

CONJUGATE FOCI.—The luminous point and its focus are interchangeable, and therefore are called *conjugate foci.*

Concave lenses increase the divergence of light rays, consequently have no *real* but a *virtual* focus. Images with a concave lens are *virtual*, *erect*, and *smaller* than the object. If the source of light be extensive, then rays from each point form their own focus, and the collection of foci forms an image, *inverted if the foci are real*, or *erect if the foci are virtual.*

The PRINCIPAL POINTS of a lens are the two points on the optical axis from which a line drawn from the first principal point to a point on the object will be parallel to a line drawn from the second principal point to a corresponding point on the image.

The PROPORTIONATE SIZE OF THE IMAGE AND OBJECT depends on their respective distances from the optical centre or the principal points.

The Aberrations of the Lens.

The image by a single lens is always indistinct because of *spherical* and *chromatic aberration.*

Spherical aberration is caused by the converging of rays from the margin of the lens, to a point *nearer* the lens than that at which the central rays meet.

Chromatic aberration is the result of the formation of different foci by rays of different color (*i. e.*, length).

Apochromatic Lenses.—These aberrations are overcome by combining two or more lenses in proper proportion—*i. e.*, concave and convex lenses made from materials having different refractive indexes.

The Photographic Camera.

This is a dark box, with a lens or system of lenses so arranged that a *real inverted image* falls on a sensitized plate in the *real focus* of the lens for the particular source of light. Decomposition of silver salts on the plate is caused by the actinic rays, and the inverted image or negative is formed. The real focus is obtained by changing the distance between the lens and a ground-glass plate, until a distinct image is observed, before introducing the sensitized plate.

The Stereopticon or Projecting Apparatus.

This is the converse of the camera: the light and inverted image are behind the lens; accordingly rays passing out diverge, and if projected on a screen at the conjugate focus in a dark room will give an erect magnified image.

The Microscope.

This is an instrument which magnifies objects through a suitable system of apochromatic lenses. In addition to the system of lenses in an opaque tube, there are arrangements for focusing and illumination of the object. According to the curve and number of the lenses, microscopes of many degrees of magnifying power have been constructed.

The Telescope.

This is constructed for viewing distant objects, and consists of an object-glass (lens) forming a *real*, *small*, and *inverted* image, which is viewed through a microscope. The *viewed* image appears *upright* through two lenses inserted between the eye-piece and the object-glass. The power of the instrument depends on the size and curve of the object-glass.

THE INTENSITY OF LIGHT.

This is measured by comparison with the light given by an English sperm candle burning 120 grains an hour, and called *candle power*.

The Rumford Photometer.

The comparative intensity may be measured by means of Rumford's photometer, consisting of an opaque rod placed a short distance from a screen with the two sources of light so arranged that each forms a shadow of the rod on the screen. The lights are then moved until the shadows are of equal intensity.

The Law of Intensity of Light.

As the light from both sources illuminates the screen proportionately when the shadows are equal, the *intensity of the lights is as the square root of their distances from the screen.*

THE CHEMICAL EFFECTS OF LIGHT.

Familiar instances out of many are: Sunlight is essential to plant-life. The green of the leaves fades and the plant dies if deprived of light. Chlorine gas and many other chemicals are made active by exposure to sun-light. Many substances—*e. g.*, silver nitrate—must be stored in dark bottles to prevent decomposition by light. The art of photography depends on the decomposition of silver compounds by light.

QUESTIONS.

What is light, and how is it transmitted?
How does red light differ from violet?
What are the effects of infra-red and actinic rays?
What waves exist in day-light, and why do different substances show difference of color in this light?
What are meant by a transparent, a translucent, and an opaque body?
What are meant by reflection, refraction, and dispersion of light?
What are the index of refraction?
What is the critical angle?
What is an optical prism?
What is a prismatic spectrum?
How may the color of the spectrum be further dispersed?
What are optical lenses, and what are the common varieties of lenses?
What are the axis, the diameter, and the optical centre of a lens?
What is meant by principal, real, virtual, and conjugate focus?
What is meant by the principal points of a lens?
What are the causes of spherical and chromatic aberration, and how are these faults overcome?
What are the essential features of the photographic camera, stereopticon, compound microscope, and telescope?
Describe a spectroscope.
What are continuous, bright-line, and absorption-spectra?
What is the cause of Fraunhofer lines?
What is double refraction?
What is meant by ordinary and extraordinary rays of light?
What is a Nicol's prism?
What is polarized light?
What is meant by optically active, dextro-rotary, and lævo-rotary substances?
What is meant by the specific rotation of a substance?
What is meant by candle power?
Mention some chemical changes produced by light.
What is meant by fluorescence and phosphorescence?

CHAPTER IV.

ELECTRICITY.

DEFINITION.—In its *earliest use*, the term electricity applied to phenomena of attraction and repulsion observed when amber (Greek *electron*) and similar substances were briskly rubbed; in its *present use*, it includes attractions and repulsions of matter very different from those originally noted, together with phenomena of light, heat, magnetism, and chemical action, all attributed to a common cause.

E THEORY OF MOTION IN ELECTRICITY.

Electricity, like heat and light, results from molecular vibration, imparted as waves to the ether, and producing undulations which in turn excite vibratory molecular motion. The action of oars on still water is an example of this mutual reaction, their stroke producing waves, capable of vibrating other oars resting on the water. *The different* **essences** *of molecular motion producing heat, light, or electricity, are not known. The* **cause** *is supposed to be a difference in the form of the undulation, heat being produced by one form, light by another, and electricity by a third. All three forms may coexist in the same body without interference with each other.*

Electricity may be regarded as a manifestation of energy consisting in undulations of the ether, and of the grosser molecules of matter. Unfortunately man possesses no sense for perceiving these waves, consequently the development of the science has been very slow.

Under normal conditions all matter is in electrical equilibrium. Whenever this equilibrium is disturbed, a tendency to its restoration is simultaneously developed. Work may therefore bê done electrically. This action of electricity is similar to the action of water pumped to an elevation, which released and descending to its former level may do mechanical work. Similarly electricity in the endeavor to restore equilibrium may be made to do work, the *quantity* of which depends on the *extent* of disturbance in the equilibrium.

THE MANIFESTATIONS OR STATES OF ELECTRICITY.

In studying electricity it is usual to divide the subject into (1) FRICTIONAL or STATIC electricity, (2) CURRENT or VOLTAIC electricity, and (3) MAGNETISM: different manifestations of the same energy.

Frictional Electricity.

If a rod of amber, glass, sealing-wax, hard rubber, or sulphur be rubbed with silk, the rod is given the power of attracting light bodies

(small pieces of paper, pith, etc.), of causing sensations similar to the touch of a cobweb, and of passing a spark with a crackling sound between the knuckle and rod after approximation: phenomena due to electricity generated by the rubbing. Certain other substances not mentioned above show similar behavior when rubbed with silk or flannel.

Conduction, Non-conduction, and Insulation.

A rod of metal rubbed with silk or flannel shows no such property unless fixed in a handle of glass, rubber, or some similar substance during the friction. The rod of metal after being electrified will differ from the rod of glass, etc., in attracting *at any point on its surface*, while the glass or similar rod only attracts *at the spot rubbed.* This shows that in metals the electricity as soon as generated spreads out over the entire mass, while in glass, etc., it remains at the point of production. *Those substances showing the latter property are called* NON-CONDUCTORS *or* INSULATORS, *while the former are called* CONDUCTORS. *A conductor supported by a non-conductor is said to be* INSULATED.

No substance is absolutely non-conducting, but a very great difference in the possession of this power exists between so-called non-conductors and good conductors. All metals, charcoal, acids, saline solutions, living animals, vegetables, water, moist earth, and stones are good conductors; while shellac, rubber, glass, resins, sulphur, wax, silk, wool, porcelain, dry paper, and dry air are poor or non-conductors. The metal bar in the hand cannot be electrified by friction, because the body too quickly conducts to the earth.

"Positive" and "Negative" Electricity.

If a pith ball be suspended by a thread from an insulated support, and approached by a *glass* rod electrified by friction with silk, the ball is attracted to, touches the rod, then falls away. If the rod be again similarly excited and approached, the ball will be repelled as long as the ball retains the electricity acquired by that contact. If, *while the ball is repelled by the glass rod*, an electrified rod of *resin* approaches, the ball will be attracted. After touching the *resin* the ball will be *attracted* by the *glass* and *repelled* by the *resin*, showing that the electricity produced by friction of resin has an opposite effect on pith balls from that which results from friction of glass.

These two forms of electricity have therefore been named RESINOUS or NEGATIVE, and VITREOUS or POSITIVE electricity (indicated by — and + signs). ELECTRIFIED or CHARGED BODIES may by experiments be shown to have either positive or negative electricity.

The Electroscope.

The nature of the charge is shown by this instrument. A **gold-leaf electroscope** consists of a glass jar, through the stopper of which

passes a metal rod terminating in the disc above, and carrying two strips of gold leaf face to face below, and two other rods or strips of metal opposite these leaves against the inside of the jar. If a *positively* electrified body touches the disc, the charge passes through the rod to the gold leaves, which, becoming *positively* electrified, repel each other. If now a second body charged with an *unknown* variety of electricity touch the disc, the leaves will further diverge if the charge is *positive*, while if *negative* they will approach each other. If the divergence causes the gold leaves to touch the metal rods on the side of the jar, the electricity of the leaf will be discharged.

When the *discharge of an electrified body* occurs it is brought to its normal condition. The discharge may occur *slowly* through the air, or *rapidly* by contact with the earth directly or through a conductor. A flash of light accompanying the discharge is called a *spark*. *Whenever one kind of electricity is generated, an equal quantity of the opposite kind is generated. In rubbing glass with silk the glass is charged with positive and the silk with negative electricity in equal quantity.*

Electrical Equilibrium.

If equal charges of positive and negative electricity are received by a conductor at the same time, it will show no trace of electricity, the two charges neutralizing each other. As has been said, in all matter electricity exists in equilibrium, and in order to show its presence it is necessary that this be disturbed. Electricity becomes manifest in restoring this equilibrium. To disturb this electric balance, *energy* in some form must be expended—*e. g.*, that of the muscles in rubbing glass with silk, of chemical action in the battery cell, or of heat in the combustion of coal in working the dynamo.

The Quantity of Electricity.

This is the normal charge possessed by any body at all times and under all circumstances, is constant for each body, and may be neutral or excited. In the latter case the positive or negative form is in excess, and the quantity of the one may be increased by removing an equal quantity of the other; but the total quantity remains unchanged, although the *quantity*, *intensity*, *tension*, or *density* of one kind is increased.

THE INDUCTION OF ELECTRICITY.

The *neutral condition* of an unelectrified body is destroyed under the influence of a electrified body. There will be attracted to the side next to the electrified body a charge, *opposite in kind* to that of the electrified body, and at the same time an equal quantity of the *other* kind is driven to the farther extremity of the neutral body. This action is

INDUCTION. When a body is connected with the earth, induction can last only while it is under the influence of the charged body, because the earth will immediately carry off the induced electricity when this influence is removed. If the neutral body, insulated and under the influence of an electrified body, be touched at the end opposite that nearest the charged body, in order to remove the electricity gathered there, the previously neutral body will be found electrified with a charge opposite in kind to that in the originally electrified body. When the original electrified body is removed, *the second body has been charged by induction.* When the pith ball is attracted by the glass rod its negative electricity is attracted to the side near the glass; on contact this *negative* electricity is discharged and the ball becomes charged with *positive* electricity by contact, and is then repelled, both bodies being charged alike.

Dielectrics are bodies through which induction occurs: air is a good dielectric.

Polarization of a dielectric is its condition of strain under the action of two opposite and adjacent electric charges. When two oppositely charged bodies approach near enough, the *limit of polarization* will be reached, and a spark passes between the bodies, as the opposite charges are neutralized through the dielectric.

If two bodies charged oppositely are connected by a *conductor*, electricity will pass *from the positive to the negative* until neutralization occurs. The former is said to have the *higher potential.* In an exchange electricity always passes from the higher to the lower potential.

THE GENERATION OF ELECTRICITY.

Static Electric Generators.

Static or *frictional electricity* is generated for practical purposes by an electric machine, consisting essentially of a plate or wheel of glass, or some similar substance, revolved by a mechanical contrivance against some material capable of producing positive electricity on the plate. This positive charge is carried from the plate by metal combs or rods to the *prime conductor;* the *negative generated on the rubbing cushions* may be collected on a conductor, but is generally carried to the ground by a chain.

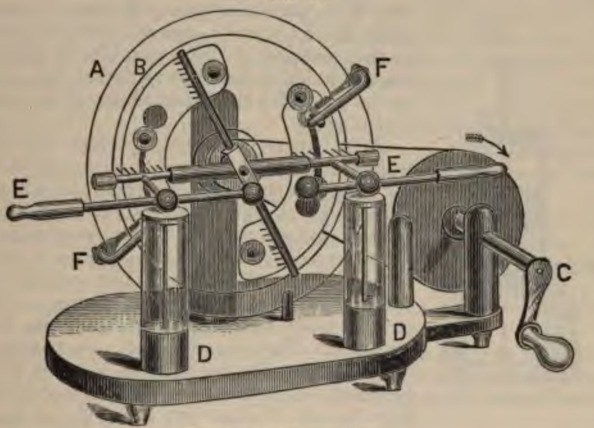
The Toepler–Holtz Machine.

This generator and its modifications act by induction much more powerfully than purely friction-generators, and are less influenced by atmospheric conditions. Fig. 12 shows its essential PARTS, namely, two plates, one *fixed* (*A*) and the other (*B*) *rapidly revolving* by the wheel and belt (*C*), *condensers* (*DD*), and *discharging rods* (*EE*), *metallic brushes* (*FF*), *two armatures* on the back of the stationary plate, and *small brass discs* on the anterior surface of the revolvable plate.

ACTION.—When this plate is revolved, its *discs* rub against the *metal-*

lic brushes. A small amount of electricity, thus developed, is carried to the *armatures* on the back of the large stationary plate. The initial charge, thus given to this plate, reacts upon the revolving plate, devel-

FIG. 12



Toepler-Holtz machine.

oping *greater charges by induction.* With this machine electricity may be applied to the body through the clothes, a fact greatly appreciated by patients.

THE CONDENSATION OF ELECTRICITY.

Electric Condensers.

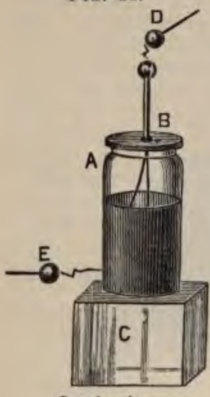
These instruments for accumulating electric charges consist essentially of two conducting surfaces separated by a non-conductor, one connected with the earth, and the other with the prime conductor of an electric machine.

The Leyden Jar.

This is a common form of condenser, and consists of a glass jar (*A*) (Fig. 13) coated on both sides nearly to the top with tin-foil, and provided with a wooden stopper (*B*), through which passes a knobbed metal rod attached to a chain which touches the foil. When the Leyden jar on an insulated stand (*C*) is placed near the prime conductor (*D*) of a working static machine, only a few sparks pass to it. If the knob

E connected with the ground is then brought near the exterior coating, a spark passes to it. For each spark thus entering the jar one spark passes to the knob *E* for some time. When the shower of sparks ceases, the machine has no further effect on the jar, which is then fully charged. Examination shows that if + electricity enters the jar, + electricity escapes. It is *not the same charge*, as is shown by the cessation of this action after a time. That the jar holds electricity even after this, is shown by the powerful spark and shock of its discharge by contact.

FIG. 13.



Leyden jar.

THE CHARGING OF A LEYDEN JAR BY INDUCTION.—The + charge entering the inner lining polarizes the dielectric, the outer foil attracts a negative charge to the surface of the glass which is bound there. The corresponding positive electricity of this surface passes to the knob connected with the earth. The negative charge reacts through the dielectric, binding a + charge on the inner foil, whose energy thus becomes potential. The inner foil may now receive a new charge repeatedly until the capacity of the jar is accumulated in a state of great molecular strain.

The amount of electricity accumulated by a Leyden jar may be increased by enlarging the jar, but very large jars are bulky.

The Leyden Battery.

This is used for large accumulations, and consists of several jars, with the *outer* coatings united by *one* and the *inner* coatings by a *second* conductor.

CONSTANT CURRENT ELECTRICITY.

This variety is generated by chemical action.

Galvanic or Voltaic Cells.

ELEMENTS AND ACTION OF THE CELL.—Zinc placed in a vessel containing dilute sulphuric acid is gradually dissolved, evolving hydrogen gas. Platinum is not acted on by dilute sulphuric acid. If plates of both metals be placed in sulphuric acid and water, and connected by a piece of wire as a conductor, no bubbles of gas are evolved on the zinc, but escape from the platinum plate. If the wire be placed near a magnetic needle, this will be turned from its course (*deflected*); when the ends of the divided wire touch the tongue, a tingling and metallic taste are perceived. All these phenomena disappear if the connection between the plates is broken, and reappear when restored, and are *manifestations of electricity* generated by the action of the acid on the metals, travel-

ling through the wire to which it imparts its properties. The zinc, platinum, and sulphuric acid may be replaced by many other materials. *Electricity will be generated whenever two solid conductors, usually in plate- or cylinder-form, are connected by a conductor, and placed in a liquid having the power of acting chemically on one of them.* Such an arrangement is called a *voltaic* or *galvanic cell;* the liquid the *electrolytic* or exciting agent; and the solids the *electrodes:* and all are called the *elements* of the cell. In a galvanic cell the electrolytic agent is split into two parts, called *ions;* and these ions, charged with + and — electricity, respectively, unload the charges on the plates. The conductor restores the electrical equilibrium. In the zinc, platinum, and sulphuric acid cell, the acid is split up into hydrogen and SO_4, the H_2 escapes from the platinum plate while the SO_4 combines with the zinc, forming zinc sulphate, which dissolves in the water.

POLARIZATION OF THE PLATES.—This is the decrease or destruction of the working of the cell by the bubbles of hydrogen collecting on the platinum plate, thus preventing the passage of electricity to the plate or setting up a countercurrent. To avoid this, a porous cup and some oxidizing agent are introduced; the latter combines the liberated hydrogen and prevents its accumulation.

Daniell Cell.—This galvanic cell consists of a glass jar containing a saturated solution of cupric sulphate, in which a copper cylinder stands; inside this is a porous cup containing dilute sulphuric acid, into which a zinc rod dips.

Leclanché Cell.—ELECTRODES, zinc and carbon in a porous cup containing manganese dioxide and small lumps of carbon; ELECTROLYTIC AGENT, saturated solution of ammonium chloride.

Grove's Cell.—ELEMENTS.—Porous cup containing platinum in nitric acid, and zinc in sulphuric acid.

Grenet or Bichromate Cell.—ELECTRODES, zinc and carbon; ELECTROLYTIC AGENT, solution of 1 lb. each of bichromate of potassium and sulphuric acid in 2½ gallons of water.

In all these cells the *zinc* plate is known as the *cathode* or *negative* electrode; the other, as the *anode* or *positive.* The end of the wire leading from each plate is also called *anode* or *positive* and *cathode* or *negative*, respectively. *When these poles are joined by a conductor, electricity passes through the wire from the + to the — electrode, and through the electrolytic agent back to the starting-point.*

Electromotive Force.

This fact shows that the cell has the power of setting electricity in motion, called ELECTROMOTIVE FORCE (E. M. F.), *tension* or *potential of the cell.* The *source* of the power is the tendency to restore electric equilibrium. The *amount* of power depends on the difference in the electrical condition of the plates. *The greater the difference the greater the E. M. F.*

This passage of electricity is the *electric current*, and its route is the *circuit*, said to be *closed* as long as the connection between the poles is intact, and *open* if the connection be broken. *When the circuit is open, the electrodes become charged with + and — electricity, respectively, which prevents further accumulation and chemical action.*

The Galvanic Battery.

This consists of two or more galvanic cells in one circuit.

The cells of a battery may be connected in either of two ways, according to the results desired.

Multiple Arc Connection.—If an *increased quantity* is required, the cells must be joined in *multiple arc—i. e., all the positive electrodes are connected by one wire and all the negative ones by another.*

Series Connection.—If *increased intensity* is required, the cells must be joined in *series—i. e., the positive electrode of one cell is connected to the negative electrode of the next.*

A single cell is also called a galvanic battery.

THE UNITS OF ELECTRIC MEASUREMENT.

The units of *intensity*, *quantity*, and *resistance* of electric currents are, respectively, the *volt*, the *ampère*, and the *ohm*, named after the three great pioneers in electricity.

The Ohm.

Conductors offer RESISTANCE to an electric current, *whose unit is the ohm—i. e.*, the amount of resistance shown by a column of mercury 106.28 millimeters long, with a base 1 millimeter square, at 0° C.

The Ampère.

The UNIT OF QUANTITY *is the ampère—i. e.*, that required to deposit 0.001118 gram of silver at the negative pole of a battery when a solution of silver nitrate is placed in the circuit.

The Volt.

This is the UNIT OF INTENSITY (*E. M. F.*)—*i. e.*, that required to maintain a current of 1 ampère through a resistance of 1 ohm.

The Law of Ohm.

The relation of the three units is expressed in *Ohm's law: The current is equal to the E. M. F. divided by the resistance.*

Practically resistance is measured by wire coils of known resistance, called *resistance coils*, and the quantity by *ampèremeters* or *ammeters*, in which a magnetic needle moves over a dial graduated in ampères. The construction of these instruments depends on the fact that when an electric current passes through a wire placed parallel above or below a magnetic needle, the latter is deflected toward a position at right angles to the wire, in degree according to the quantity of current.

A current of 1 ampère is too strong for medical use; hence the **milli-amperemeter,** *measuring* $\frac{1}{1000}$ *ampère, has been devised. Current electricity has great quantity and low tension. Frictional electricity has smaller quantity but higher tension.*

Magneto-electric Machines.

These generate electricity by magnetism. Whenever metallic wires are passed through a magnetic field, temporary electric currents are set up in them.

ALTERNATING CURRENT ELECTRICITY.

If two coils of insulated wire with soft-iron cores (armatures) be made to revolve rapidly in front of or between the poles of a horseshoe magnet, the cores will be **magnetized** *while opposite one of the poles of the magnet, and* **demagnetized** *when equidistant from the poles, and* **remagnetized** *when opposite the other pole.* This magnetization and demagnetization induces currents in the wire, running in opposite directions, and alternating as the polarity of the cores changes with each half-revolution. Such currents are called *alternating currents.*

The Dynamo-electric Machine (Dynamo).

In this machine powerful electro-magnets and armatures with many coils are used to generate very powerful currents.

The COMMUTATOR collects alternating currents, changing them to a continuous current.

Electric Motors.

These are *reversed dynamos.* Electric currents convert iron into electromagnets, which by alternate attraction and repulsion cause armatures to revolve.

The method of constructing an *electromagnet* has been mentioned. Their strength depends on the length of wire around the core and the strength of current passing through the wire, therefore very powerful electromagnets may be made. *The position of the north and south poles of an electromagnet depends on the direction of the current. A reversion of the latter causes a reversion of polarity in the magnet.*

VOLTAIC INDUCTION.

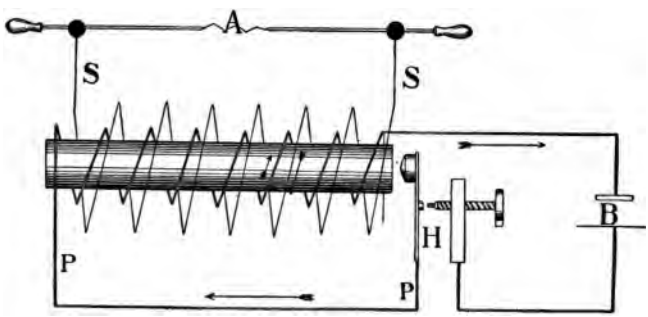
Like static electricity and magnetism, current electricity passing through one wire can induce currents in adjacent wires.

THE RUHMKORF INDUCTION COIL.

This instrument induces currents of *high tension* (E. M. F.) from battery currents of *low tension*.

The structure of the coil is shown diagrammatically in Fig. 14: a *core of soft iron*, wound with relatively *coarse wire* (*P*) connected with a *battery* (*B*); a *secondary coil* (*S*), formed by winding a *much finer* and *longer wire* around a *hollow cylinder*, is arranged to slip over the first or

FIG. 14.



Induction coil.

primary coil, but the two are separate and distinct. With the circuit *open*, a current passes through the primary without influencing the secondary coil, but as soon as the circuit is *closed* instantaneously a current is induced in the secondary wire in *one* direction, and on *opening* the circuit in the *opposite* direction. The automatic *opening* and *closing of the circuit* is accomplished by a movable hammer (*vibrator*) (*H*) held by a spring to close the circuit when the core will be magnetized. The electromagnet so formed attracts and raises the hammer, thereby opening the circuit. The core is now demagnetized by cessation of the current, and the hammer returns to its original position and closes the circuit. The action is renewed, and continues to rapidly open and close the primary circuit as long as the current flows, generating in the secondary wire that current known as *secondary*, *induced*, *interrupted*, or *faradic current*.

The current generated at the closure of the circuit is the *make*, and that at the opening is the *break* current. The effectiveness of an induction coil increases with the length of its wire, which in large instruments

is over fifty miles. With such machines sparks will pass several feet between the terminals of the secondary coil at *A*.

THE CONVERSION OF ELECTRICAL ENERGY.

Electric Heaters.

Resistance to the passage of electricity changes the electrical energy into heat, which is utilized in cars and buildings, stoves for cooking, vulcanizers used in dental manipulations, and electric furnaces. The degree of heat depends on the structure of the apparatus and the strength of current used. In the electric furnace strong currents passing through carbon meet with sufficient resistance to give rise to temperatures unattainable by any other means.

Electric Lights.

The Electric Arc Lamp.—Here a current of high tension (1000–3000 volts) passes between terminals consisting of pencils of carbon, held in place by an automatic contrivance. The carbon is heated to luminosity, and the current passing from the + to the — pole carries particles of carbon which are burned or deposited on the negative pole.

The Incandescent Lamp.—Here a current of 50–120 volts heats to a white heat a filament of carbon contained in a glass vacuum-bulb.

Electrolysis.

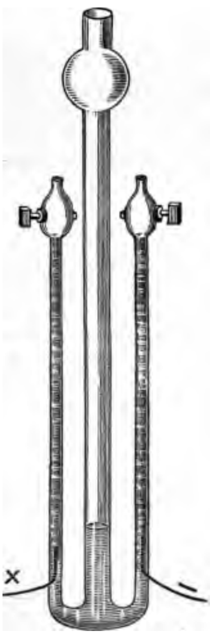
This is chemical decomposition—*i. e.*, the separation of matter into its atomic components by electricity, and is a conversion of electrical into chemical energy.

Fig. 15 represents an apparatus for decomposing water by electricity, and collecting and measuring the component gases; consisting of three glass tubes opening into one another. In each outer tube platinum foil is connected with a platinum wire fused into the glass and passing to the outside of the tube. The tubes are filled with water acidified with sulphuric acid, and the platinum wires connected with the wires from a battery, then oxygen rises from the positive pole and hydrogen from the negative pole.

ELECTROLYTES are bodies undergoing electrolysis. *Ions* are the decomposition-products, those collected at the anode or + pole are *anions*, and those at the cathode or — pole *cations*. *Anions* are said to be *electronegative*, and *cations*, *electropositive*. Hydrogen, metals, bases, and basic radicals are cations. Oxygen, the halogens, acids, and acid radicals are anions. Pure water cannot be decomposed by electricity, although the constituents of pure water may be obtained by electrolysis as described above. The result is believed to be due to decomposition of the sulphuric acid. It is supposed that the sulphuric acid (H_2SO_4) splits into the ions of H_2 and SO_4, and that the cation carries its charge to the negative

pole and is there discharged, while the anion is discharged at the positive pole. The hydrogen is liberated at the negative pole as gas bubbles. The SO_4 after discharge acts chemically on a molecule of water (H_2O), combining with the hydrogen to form sulphuric acid, oxygen being set free.

FIG. 15.



Electrolysis of water.

A salt in solution will be split into its component base and acid by the passage of an electric current through the solution. If the base be a metal—*e. g.*, copper, nickel, silver, gold—it will be deposited on the negative pole. The processes of electrotyping and electroplating depend on this property. Electrolysis is used in analyses, in the separating of metals from their ores, and in electric furnaces. In the latter the heat and decomposing force come from the electric current.

All liquids which electrolyze are conductors, and contain some substance which suffers *dissociation* (*separation*) into ions, thus furnishing carriers for the current. It is supposed that an electrolyte in passing into dilute solution dissociates into *two* ions—*e. g.*, common salt in water separates into sodium and chlorine. When the solution is concentrated by evaporation the ions combine into sodium chloride.

Electric sparks are luminous effects of electric currents of sufficiently high E. M. F. passing through gases, rendering the gas luminous, and frequently exerting chemical action. If the discharge is in a glass globe, and the gas be gradually removed from it, the sparks cease at a reduced pressure, and the interior of the globe becomes luminous, the nature of the luminosity depending on the gas in the globe.

Rontgen or X-rays.

These are rays formed by electric sparks passing through a vacuum-tube provided with a metallic disc cathode and a straight wire anode, and manifested by a pale purplish beam radiating from the face of the disc. This *cathode ray* consists of negatively charged portions of matter moving with great velocity. During the generation of the cathode rays the Röntgen or *X*-rays are generated, which are invisible and have the property of passing through substances opaque to ordinary light, of affecting photographic plates like ordinary light, of causing many substances to become fluorescent, and of acting physiologically on various parts of the animal body.

The Fluoroscope.

This is an instrument for observing with the eye the effect of *X*-rays, depending on the property of the rays to cause fluorescence of certain substances, and is usually constructed of a box closed at one end by a screen of some fluorescent substance (*calcium* or *barium tungstate*), and provided at the other end with cushions to fit closely to the face of the observer. When certain opaque objects, as a hand or arm, are placed in the path of *X*-rays shadows are cast on the screen.

Radio-active substances have the property of emitting rays which act like Röntgen rays. Examples are: actinium, polonium, and radium, substances found in minute quantities in ores of uranium.

CHAPTER V.

MAGNETISM.

DEFINITION.—Magnetism is that peculiar property in some substances, called MAGNETS, of attracting or repelling other substances under certain conditions.

NATURAL MAGNETS.

LOADSTONE OR MAGNETITE, a native iron ore (*ferrous-ferric oxide*), naturally attracts bits of iron and steel, and a fragment of this ore when suspended will tend to point north and south. These properties may be *acquired* by iron, steel, and a few other metals (cobalt and nickel). *Loadstone is a* NATURAL MAGNET, *while bars of metal which have acquired magnetic power are called* ARTIFICIAL MAGNETS.

ARTIFICIAL MAGNETS.

These are made by several processes:

1. BY SIMPLE TOUCH: the bar to be magnetized is stroked several times in *one direction from end to end* by another magnet.

2. BY DOUBLE TOUCH: the bar is stroked *from the middle to one end* and *back to the middle* with the *opposite* poles of two other magnets *separated by a piece of wood.*

3. BY SEPARATE TOUCH: the *opposite* poles of two other magnets are placed at the *middle of the bar* and simultaneously moved away from each other to the ends.

4. BY MAGNETIC INDUCTION: a soft-iron bar near a magnet will attain magnetic properties, which it will retain only as long as it is under the influence of the magnet.

5, BY ELECTROMAGNETIC INDUCTION: a bar of metal surrounded with a coil of wire through which a current is passing becomes magnetic (*electromagnetic*).

THE MAGNETIC POLES AND MERIDIAN.

The earth may be considered as a huge magnet, whose *magnetic poles* are *near*, but not *at* the *geographical poles*. Therefore the earth causes a suspended magnetic needle to occupy a plane, called the *magnetic meridian*, nearly north and south, and to incline one pole toward the earth. The end of the magnet pointing north is called the **north pole**, the other end the **south pole**. On bringing the north pole of another magnet near one pole of a suspended magnet, it will be observed that with magnetism, as with electricity, like poles repel and unlike poles attract.

MAGNETIC ATTRACTION AND REPULSION.

MAGNETIC SUBSTANCES are attracted by magnets, but are not themselves magnetic.

PARAMAGNETIC BODIES (iron, steel, cobalt, nickel, and platinum) are more or less easily attracted by magnets.

DIAMAGNETIC BODIES are repelled by a magnet (bismuth, tin, zinc, antimony, mercury, lead, silver, copper, gold, and arsenic).

When a magnetized bar is dipped into iron-filings, these adhere at its ends, called *the poles of the magnet*, and not at its centre, called *the neutral line* or *equator*. A straight line joining the poles is the *axis*.

THE MOLECULAR THEORY OF MAGNETISM.

If a magnet be halved through its equator, each half will be a complete magnet, having two poles and an equator, and each, in division, will form two smaller, complete magnets, no matter how many times the division is made. *This and some other reasons have led to the theory that each molecule of a magnet is a magnet itself.* In ordinary iron the molecular magnets are not arranged systematically, but in magnetized iron all north poles point in one direction, the south poles in the other. The maximum attraction of the end is all these individual forces combined.

THE THEORY OF PARALLEL MAGNETIC CURRENTS.

Each molecule is assumed to be traversed by a closed electric circuit, whose currents become parallel when the iron is magnetized; thus magnetism is a manifestation of electrical energy. The theory derives support from the fact that a spiral conductor (called a SOLENOID) traversed by a current behaves as a magnet in all respects.

MAGNETIC SATURATION.

A magnet is said to be saturated when it has been magnetized to the *highest* degree possible, supposedly when all north poles of the molecules point in one direction.

THE COERCIVE FORCE OF MAGNETISM.

This condition of a substance retards the molecular arrangement of magnet-formation and retains this arrangement after magnetization. Steel has high coercive force, and therefore is difficult to magnetize, but remains so permanently. Soft-iron magnets are the reverse, through low coercive force. Artificial magnets are usually shaped as bars or horseshoes. The bar of soft iron connecting the ends of a horseshoe-magnet is called an *armature*, and prevents loss of power. COMPOUND MAGNETS, MAGNETIC BATTERIES, or MAGNETIC MAGAZINES, are the combination of several magnets (usually horseshoe) with like poles united.

The ASTATIC NEEDLE or MAGNET is two magnets, placed parallel, on a common axis, with unlike poles suprajacent. Such needles are not affected by the earth's magnetism, but remain in whatever direction they are placed.

THE LINES AND FIELD OF MAGNETIC FORCE.

THE MAGNETIC FORCE-LINES.—When a magnet is covered with a sheet of glass or cardboard and iron-filings sprinkled on the cover, the filings arrange themselves in lines radiating from each pole, and curving from pole to pole, called *lines of magnetic force*, and representing the resultant of the combined action of the two poles.

THE MAGNETIC FIELD is the space surrounding a magnet as far as its influence is exerted. A magnet suspended therein follows the lines of magnetic force, lying tangent to the magnetic curve.

QUESTIONS.

What is meant by electricity?

What is the difference between static and voltaic electricity?

What is meant by a conductor and an insulator?

What is meant by positive and negative electricity?

Describe the electroscope and its workings.

What is the difference between quantity and tension of electricity?

What is induction?

What is a dielectric?

What is meant by polarization of a dielectric?

Explain the action of a Troepler-Holtz machine.

What is an electric condenser?

Describe the structure and method of charging a Leyden jar.

What is magnetism, and what is meant by natural and artificial magnets, and by a solenoid?

What is meant by the north and south poles of a magnet?

What is meant by magnetic, paramagnetic, and diamagnetic bodies?

What is meant by the equator of a magnet, magnetic field, lines of magnetic force, and magnetic magazine?

What is coercive force?

Describe an astatic needle.

Describe the construction and workings of a galvanic cell.

What is meant by the cathode and anode poles?

Define electric current and circuit.

What is meant by the E. M. F. of a cell?

What is the difference between an open and a closed circuit?

How are cells joined in multiple arc and in series, and what is the result of each method on the circuit?

What are the three principal units of measure used for electric currents, and how are they determined?

Describe the structure and workings of a magneto-electric machine, and a dynamo.

What is a commutator?

What is an electric motor?

Describe the induction coil.

Describe the electric furnace, arc light, and incandescent light.

What is meant by electrolysis, electrolytes, and ions?

Explain the theory of electrolysis.

What are the properties of Röntgen or *X*-rays, and how are they generated?

What is the fluoroscope?

What is meant by radio-active substances?

PART II.

INORGANIC CHEMISTRY.

CHAPTER I.

DEFINITION.—Chemistry is the department of science treating of the composition of substances, the changes in this composition from any cause, and the laws governing the changes.

THE FUNDAMENTAL PRINCIPLES OF CHEMISTRY.

Chemical Divisibility.

A piece of ice heated in a glass tube will melt, form steam, and disappear. Red oxide of mercury, heated, similarly gradually disappears. If these substances be heated in a tube connected with apparatus for collecting and cooling the gases formed, the gas from ice may be reconverted into water and ice, if the temperature be sufficiently lowered. From the red oxide of mercury a silvery liquid is formed and a colorless gas which cannot be reconverted by *physical* action. On the red oxide of mercury the heat has therefore produced an effect entirely different from that on the ice, because it first produced its ordinary increase of molecular motion and volume, after which the oxide decomposed, forming two substances, each different from itself: one, a silvery heavy liquid metal (*mercury*), the other, a colorless gas, supporting combustion vigorously (*oxygen*).

Elements and Compounds.

DEFINITION.—Mercury and oxygen resist all attempts to convert them into simpler forms, issue unchanged from all varieties of reactions with this intent, and are not decomposed by the most powerful influences.

All substances which, like oxygen and mercury, cannot be converted into simpler substances by any known means, are called *elements*, while those which may be separated into simpler substances with new properties, are called *compounds*.

Molecules and Atoms.

DEFINITION.—*Molecules* are the smallest particles of matter existing free. All matter consists of molecules. From red oxide of mercury mercury and oxygen are obtained, each consisting of molecules. As the molecules of red oxide of mercury contain both mercury and oxygen, and as all molecules are of the same size (*Avogadro's law*), the molecules of the mercury must be made of particles of oxygen and mercury, each smaller than a molecule.

Atoms are those particles of elements uniting to form molecules. *Atoms* are also defined as the smallest *particles* of an element capable of entering into combination.

Elements have, therefore, the atoms of their molecules alike.

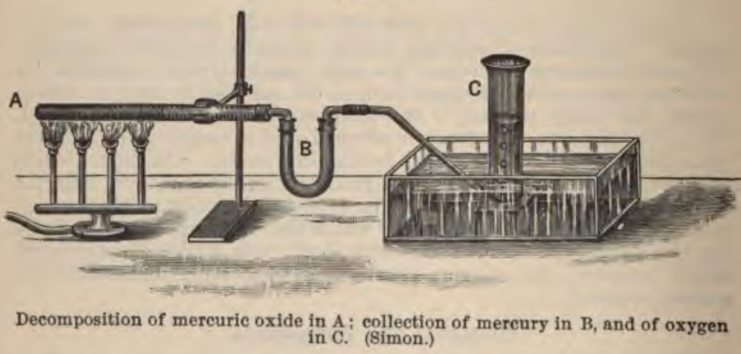
Compounds are substances, the molecules *of which contain atoms of different kinds.*

The total number of known elements is seventy-six. By combining in different proportions they form the varieties of matter in existence.

Chemical Affinity.

DEFINITION.—If the red oxide of mercury be decomposed in apparatus like that in Fig. 16. the oxygen formed collects in the jar *C*, forc-

FIG. 16.



Decomposition of mercuric oxide in A; collection of mercury in B, and of oxygen in C. (Simon.)

ing out the water. A piece of phosphorus, another element, introduced into a jar thus filled with oxygen, will ignite, burn, and form a

white powder, proved to be a compound of phosphorus and oxygen. If a piece of platinum be introduced into the jar, no change is noticed. For some cause phosphorus and oxygen combine when brought together, while platinum and oxygen do not. The cause is a form of *attraction* called *chemical force* or *chemical affinity*, by which atoms of different elements form new substances, and bodies possessing this power of uniting into new substances have *affinity* for one another. When the atoms are all of one kind they combine into molecules—*e. g.*, *hydrogen gas is a collection not of atoms, but of molecules of two atoms each.*

THE RELATION OF CHEMICAL TO THE OTHER FORMS OF ATTRACTION.—*Chemical affinity* is the attraction of atoms for each other, uniting them into molecules.

Cohesion unites the constituent molecules of a substance.

Adhesion keeps the molecules of adjacent surfaces of different substances in contact.

Gravitation is the attraction of masses for each other.

Atomic and Molecular Weight.

DEFINITIONS.—All matter has weight, and molecules and atoms have their proportion of the whole.

Atomic weight is a comparative weight, and the standard is the atom of least weight. *Hydrogen has the lightest atom, and is therefore said to have an atomic weight of unity, with which the weights of all other elements are compared.*

Molecular weight is the combined weights of all the atoms of a molecule.

The weighing of individual molecules and atoms being impossible, their weight must be determined by indirect methods. Of these, the more important are the following.

Methods of Determination of Atomic Weight.

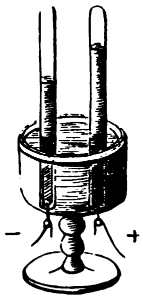
I. BY CHEMICAL DECOMPOSITION.

By Electrolysis.—A weighed quantity of acidified water is placed in an apparatus for electrolysis like that shown in Fig. 17. The tubes containing the poles are graduated, and completely filled with water with their open ends below the surface of the water in the jar. The apparatus is connected with a battery and electrolysis begun.

Hydrogen and oxygen form at the poles, rise, and force the water from the tubes. When a sufficient quantity of the gases has been formed, the electrolysis is checked and the volume of each gas is read from the graduations. The volume of hydrogen is always twice that of the oxygen. The gas in each tube is now weighed, the oxygen weighing 8 times as much as the hydrogen. As the volume of hydrogen is

twice that of the oxygen, oxygen must weigh 16 times as much as an equal volume of hydrogen. *According to the law of Avogadro, equal volumes of hydrogen and oxygen contain the same number of molecules, therefore oxygen is 16 times as heavy as hydrogen. The atomic weight of oxygen is consequently 16.*

FIG. 17.



By Substitution.—Hydrogen is liberated from a substance of known molecular weight by the element whose atomic weight is to be determined—*e. g.*: A molecule of hydrochloric acid consists of 1 atom of H, weight 1, and 1 atom of chlorine, weighing 35.4; therefore, its molecular weight is 36.4. In this acid, hydrogen may be replaced by some other element (sodium), whose atomic weight is then determined by comparing the weight of the metal combined with the weight of hydrogen liberated. If 46 grams of sodium liberate 2 grams of H from the acid, then the atomic weight of sodium is 23. This method does not show whether 1 atom of sodium replaced 1 atom of hydrogen, or more than 1, consequently there is doubt as to the correctness of the conclusion, because if 1 atom of sodium replaced 2 of hydrogen, then the atomic weight of sodium would be 46.

II. BY THE SPECIFIC GRAVITY OF GASES.

This method is not unlike the comparison of oxygen and hydrogen just described under Electrolysis. When equal volumes of a gas of unknown molecular weight and of hydrogen are weighed, a comparison of the weights will give the *molecular* weight of the gas examined—*e. g.*, if a flask of hydrogen weighs 2 grams, and the same flask of another gas weighs 64, then its *molecular* weight is 64. If it be an *elementary* gas, containing 2 atoms to each molecule, its *atomic* weight is 32.

When the element is easily volatilized this method is very simple. When it is not volatile the *molecular* weight of some gaseous product of it is determined, and from this the *atomic* weight of the element is calculated.

III. BY SPECIFIC HEAT.

It has been proved that the atomic weight is *inversely proportional* to the specific heat of an element. The product of the atomic weight multiplied by the specific heat is a constant quantity, shown by the table below to be about 6.5. These facts were formulated as a law by Dulong and Petit in 1819.

The Law of Dulong and Petit.—*The atoms of all elements have exactly the same capacity for heat.*

TABLE SHOWING THE CONSTANT PRODUCT OF SPECIFIC HEAT INTO ATOMIC WEIGHT OF ELEMENTS.

Element.	Specific Heat.	Atomic Weight.	Product.
Boron	0.5910	11	6.5
Carbon	0.5416	12	6.49
Aluminum	0.2407	27	6.5
Potassium	0.1667	39	6.5
Copper	0.1008	65	6.58
Iodine	0.0513	126.5	6.48
Lead	0.0315	206.4	6.5

The law of Dulong and Petit may be proved by taking 39 grams of potassium, 6.5 of copper, 206.4 of lead, and exposing them to the *same quantity* of heat, when the temperature of each rises equally. As these weights are the *atomic* weights of the elements, the same number of atoms of each is exposed to the heat, consequently the quantity of heat required to raise the temperature of 1 atom of potassium 1 degree will raise that of 1 atom of lead 1 degree, and so on. This is true of any other series of elements.

By determining the specific heat of an element, and dividing this number into 6.5, *the atomic weight of that element is determined.*

Methods of Determination of Molecular Weight.

The molecular weight of a compound is determined by methods somewhat similar to those for atomic weights.

I. BY CHEMICAL ANALYSIS.

The percentage composition of a substance is calculated, and from it the molecular weights determined—*e. g.*, a substance on analysis shows 46 per cent. of sodium, 12 per cent. of carbon, and 48 per cent. of oxygen. The molecule, by inference, consists of *two* atoms of sodium, *one* of carbon, and *three* of oxygen, because the atomic weights of these elements are 23, 12, and 16, respectively; hence the *molecular* weight will be 85, and the *composition*, expressed by the formula Na_2CO_3. The same percentage composition might also represent a compound having the formula $Na_4C_2O_6$. Many substances, as C_2H_4 and C_4H_8, have the *same percentage* composition, and analytic methods are insufficient to indicate the number of atoms in the molecule. *When the substance is volatile the application of Avogadro's law will avoid this difficulty.* A volume of C_4H_8 will weigh 56 grams, an equal volume of C_2H_4 will weigh 28 grams, and an equal volume of hydrogen will weigh 2 grams.

II. By the Fall of the Freezing-point (Raoult's Method).

The lowering of the freezing-point of a solution is proportionate to the number of molecules in solution. This fact is taken advantage of in determining molecular weight—*e. g.*, the freezing-point of water is 0° C. If a 5 per cent. watery solution of a salt of *known molecular* weight freezes at —1° C., then the 5 per cent. of salt has lowered the freezing-point 1 degree C. If now a salt of *unknown molecular* weight be dissolved in water sufficiently to lower the freezing-point 1 degree C., then the quantity of salt dissolved in the latter case will contain the same number of molecules as the 5 per cent. of salt in the first case. Suppose 12 per cent. of the unknown salt be required to lower the freezing-point 1 degree C., then the molecular weights of the two substances will be to each other as 5 : 12.

This method is not applicable in all cases, the two substances must have no chemical action on the solvent, and must be members of the same class of substances.

CHEMICAL DESIGNATION.

Chemical Symbols.

Letters or groups of letters are used for conveniently designating elements, and are generally the first or the first and second letters of the *Latin* name of the element—*e. g.*, O is the symbol of oxygen, H of hydrogen, Hg of mercury (Hydrargyrum), Ca of calcium, and C of carbon.

Chemical Formulas.

Collections of symbols are used to represent the constituent atoms of a molecule and the proportion of each atom. When a formula consists of symbols of elements only, as HgO, it indicates that the molecule of oxide of mercury consists of 1 atom of mercury and 1 atom of oxygen, or 200 parts by weight of mercury and 16 parts by weight of oxygen.

The figure placed *below* and to the *right* of a symbol in a formula indicates the number of atoms of the element represented which occur in the molecule—*e. g.*, H_2O indicates that a molecule of water consists of *two* atoms of hydrogen and *one* of oxygen. The figure placed on the *same* line with, and to the *left*, of a molecule indicates the number of molecules entering into the compound—*e. g.*, Ca 2(OH) indicates the molecule of calcium hydroxide to consist of *one* atom of calcium in combination with *two* molecules of a compound containing one atom each of oxygen and hydrogen.

Molecular formula is the term describing the above formulas. They show simply the different atoms and the proportion of each in the molecule.

THE LAWS OF CHEMICAL COMBINATION.

General Principles.

DEFINITION.—*Chemical combination* is the union of two or more substances through chemical affinity, into a new substance, differing in properties from any ingredient. The combining substances lose their distinct properties.

Mechanical mixture and chemical combination differ fundamentally, as best shown by an example:

If iron-filings and powdered sulphur be rubbed together in a mortar, a greenish-yellow powder is obtained. A few grains of this under the microscope will show the separate particles of iron and sulphur, and a magnet will remove all the iron from the powder and leave the sulphur unchanged. The greenish-yellow powder is a *mechanical mixture* or a *mixture of molecular clusters.*

If this mixture be heated until the sulphur melts, the elements will form sulphide of iron, a substance differing in properties from both ingredients, their characteristics having disappeared. This process is therefore a *chemical combination.*

Molecular Mixtures.—When two substances, as two metals, are fused together, or when two gases or liquids are mixed together, or when a solid is dissolved in a liquid, the individual masses can no longer be recognized by the microscope. These mixtures are called *molecular mixtures.*

The dissociation of electrolytes, on dissolving, has been mentioned (see page 56). This dissociation is not a chemical change.

A true chemical combination is not a mixing of molecules, but is a joining together of atoms into the molecules of a new substance.

The Law of Constancy of Composition.

The same substance always contains the same elements in the same proportion.

The proportion in which the atoms unite into the molecule of a definite chemical compound is invariable. In the combination of sulphur and iron into *sulphide of iron* 32 parts *by weight* of sulphur and 56 of iron unite to form 88 of sulphide of iron. These figures represent the atomic weights of sulphur and iron, and if the elements are mixed in any other proportion the excess of one will remain uncombined and unchanged—*e. g.*, if 100 pounds of iron and 32 of sulphur are heated together, 88 pounds of sulphide of iron and 44 pounds of unchanged iron will result. The fact that the proportions in which the atoms unite to form a molecule of a definite chemical compound are invariable, was the first law ever recognized in chemistry.

The Law of Multiple Proportions.

One element may unite with varying quantities of another into different compounds, but the varying quantities of the second element combining with a fixed quantity of the first bear a simple ratio to each other—*e. g.*, in the combinations of nitrogen and oxygen, into N_2O, N_2O_2, N_2O_3, N_2O_4, N_2O_5, the *varying* quantities of oxygen combining with a *fixed* quantity of nitrogen (*two atoms*) are to each other as 1 : 2, 1 : 3, 2 : 3, 3 : 4, 4 : 5, etc.

Dalton's atomic theory finds its chief support in the foregoing law, Dalton arguing that the only possible explanation of the facts cited lies in the assumption that elements are made up of particles of definite weight, which he called atoms.

The Law of Chemical Combination by Volumes. (Gay-Lussac.)

When two or more gases combine chemically into a gaseous compound, the volumes of the individual constituents bear a simple ratio to the volume of the product.

If the combining volumes of the different gases be weighed, a simple ratio between these weights and their atomic or molecular weights will be found to exist—*e. g.*, 1 volume of hydrogen and 1 volume of chlorine unite to form 2 volumes of hydrochloric acid (HCl). The proportion of H to Cl is as 1 : 1. If the volume of hydrogen weighs 1 gram, the equal volume of Cl weighs 35.4 grams—*i. e.*, the atomic weights of the elements. Under constant pressure and temperature Gay-Lussac observed the following: 1 liter of hydrogen combines with 1 liter of chlorine to form 2 liters of gaseous hydrochloric acid (molecular weight, 36.4); 2 liters of hydrogen and 1 of oxygen combine to form 2 liters of gaseous water (molecular weight, 18); 1 liter of hydrogen, 1 of nitrogen, and 3 of oxygen form 2 liters of gaseous nitric acid. Similar observations including all combinations of gaseous elements into gaseous compounds, and some other reasons, led to the adoption of the *law of Avogadro* (see page 19).

The Law of Equivalents, Valence, or Quantivalence.

"*When one element replaces another element in a combination, the replacement always takes place in fixed proportions.*"

DEFINITIONS.—*Chemical equivalents* or *equivalent numbers* are the relative quantities in which the elements replace one another.

Equivalent elements are those in which *one* atom of the one replaces *one* atom of the other element.

Atomicity, *quantivalence*, or *valence*, is the peculiar attraction possessed by the atoms of certain elements, by which they have the power of replacing two or more atoms of others.

The *unit of measure* of this power is the valence of the hydrogen atom.

Univalent elements are those *one* of whose atoms replaces *one* atom of hydrogen in a combination, or combines with *one* atom of hydrogen into a *saturated* compound.

Bivalent, trivalent, quadrivalent, quinquivalent, etc., elements are those *one* of whose atoms replaces, *respectively, two, three, four, five*, etc., hydrogen atoms. *Monads, diads, triads, tetrads, pentads*, etc., are synonyms for these terms.

Valence is often indicated by dots or numbers placed *above* and to the *left* of the chemical symbol—*e. g.*, $\overset{\cdot\cdot\cdot\cdot}{C}$, C^{iv}, indicate that carbon is quadrivalent.

BONDS OF AFFINITY.

Arithmetically *univalent* atoms have *one; bivalent, two*, etc., bonds of affinity or points of attraction to other atoms. Bonds are graphically represented by lines between one symbol and another—*e. g.*,

$$\text{H—},\ \text{—S—},\ \text{—}\overset{|}{\text{Bi}}\text{—},\ \text{—}\underset{|}{\overset{|}{\text{C}}}\text{—}.$$

They are only symbolical, and express the view of a definite power of atomic combination.

SATISFACTION OF THE BONDS is the full combination of atoms, and is expressed *graphically* as below.

GRAPHIC, STRUCTURAL, OR CONSTITUTIONAL FORMULAS.

These diagrammatically represent this satisfaction of bonds—*e. g.*:

$$\text{H—S—H or S}{<}^{\text{H}}_{\text{H}};\quad \text{H—}\underset{\underset{\text{H}}{|}}{\overset{\overset{\text{H}}{|}}{\text{C}}}\text{—H or C}\begin{array}{l}\diagup\text{H}\\ {<}\text{H}\\ {<}\text{H}\\ \diagdown\text{H}\end{array}.$$

Graphic or structural formulas do not represent the actual position of the atoms in space, but express the chemical relations of the substance.

By means of them chemical changes are made more intelligible, and methods of preparation of new compounds often foreseen before actual discovery. Molecular arrangement cannot be rendered *visible*, but can be *determined* by study of physical and chemical properties, combinations, and decompositions of substances under certain conditions.

Certain elements do not show the same valence under all conditions —*e. g., two* atoms of phosphorus can unite with either *three* or *five* atoms of oxygen into the saturated compounds P_2O_3 and P_2O_5. As oxygen is bivalent, therefore the phosphorus acts as a trivalent and pentavalent, respectively. Similar action is shown by several other atoms, and may

be explained only by assuming that one or more bonds of an atom unite with other bonds of the *same* atom, as is graphically shown by:

$$\begin{array}{ccc} \begin{array}{l} <P=O \\ \quad\;\; >O \\ <P=O \\ \quad P_2O_3 \end{array} & \qquad\qquad & \begin{array}{l} O=P=O \\ \qquad\quad >O \\ O=P=O \\ \quad\;\; P_2O_5 \end{array} \end{array}$$

Whenever the atom of an element shows variations in valence, the valence increases or diminishes by the fixed quantity *two*, which is a fact essential to the above explanation—*e. g.*, *chlorine*, usually a *monad*, may have a valence of iii, v, or vii; *sulphur*, usually *bivalent*, may be iv or vi valent.

Artiads and Perisads.

Definition.—*Artiads* are atoms whose valence is expressed in *even*, and *perisads* in *uneven* numbers.

Unsaturated Compounds and Radicals.

Definition.—*Unsaturated compounds* are formed when atoms of elements of different valences are united, leaving affinities unsaturated, and therefore combine with other elements with great ease.

Radicals are unsaturated compounds, capable of entering and leaving combinations without being changed.

QUESTIONS.

What is the difference between the action of heat on ice and on red oxide of mercury?

What is an element, and how do elements differ from compounds?

What is an atom, and what is a molecule?

What is chemistry?

What is chemical affinity?

How can the existence of affinity be shown?

How does affinity differ from cohesion, adhesion, and gravitation?

What are meant by atomic weight and molecular weight?

What atom is the unit of comparison in atomic weights?

Describe three methods of determining atomic weights.

Why are chemical means insufficient for determining atomic weights?

Explain the law of Dulong and Petit.

If the specific heat of an element is 0.9408, what is its atomic weight?

If 224 grams of an element replace 8 grams of hydrogen in hydrochloric acid, what is its atomic weight?

What are chemical symbols and chemical formulas?

What is indicated by the following formulæ: FeO, $CuCl_2$, $MgBr_2$, PbS?

How do chemical combinations differ from mechanical mixtures and molecular mixtures?

State the law of constancy of composition.

Explain the law of multiple proportions.

State and explain the law of chemical combination by volume.
Explain valence.
What is a graphic formula?
What is an artiad, and what is a perisad?
What is the structural nature of an unsaturated compound?
What is a radical?
What is the characteristic behavior of a radical?

CHAPTER II.

THE ELEMENTS AND CHEMICAL COMPOUNDS.

THE ELEMENTS.

Practical Classification.

I. Universal Importance.—Only about one-fourth of the total number (77) of elements unite to form the greater part of the earth, atmosphere, animal and vegetable matter.

II. Commercial Importance.—Other elements, not in large quantities or important in animal or vegetable matter, are used in some form or other in the arts.

III. Scientific Importance.—Elements not in the above classes are of scientific interest, and exist in minute or problematical quantities.

Table of the Elements (after Simon).

I. Universally Important.	II. Commercially Important.	III. Scientifically Important.
Aluminum.	Antimony.	Beryllium.
Calcium.	Arsenic.	Cæsium.
Carbon.	Barium.	Columbium.
Chlorine.	Bismuth.	Didymium.
Hydrogen.	Boron.	Erbium.
Iron.	Bromine.	Gadolinium.
Magnesium.	Cadmiun.	Gallium.
Nitrogen.	Cerium.	Germanium.
Oxygen.	Chromium.	Indium.
Phosphorus.	Cobalt.	Iridium.
Potassium.	Copper.	Krypton.
Silicon.	Fluorine.	Lanthanum.
Sodium.	Gold.	Neodymium.

Sulphur.	Iodine.	Neon.
	Lead.	Osmium.
	Lithium.	Palladium.
	Manganese.	Praseodymium.
	Mercury.	Rhodium.
	Molybdenum.	Rubidium.
	Nickel.	Ruthenium.
	Platinum.	Samanum.
	Silver.	Scandium.
	Strontium.	Selerium.
	Tin.	Tantalum.
	Zinc.	Tellurium.
		Terbium.
		Tetanium.
		Thallium.
		Thorium.
		Thulium.
		Tungsten.
		Uranium.
		Vanadium.
		Xenon.
		Ytterbium.
		Yttrium.
		Zirconium.

Chemical Classification.

Elements are best classified by dividing them into two groups the METALS and the NON-METALS.

The Metals.

PROPERTIES.—Metals, fifty-eight in number, possess opacity, peculiar brightness (*metallic lustre*), conduct heat and electricity well, combine with oxygen into compounds with basic properties, and replace hydrogen in acids forming salts.

The Non-metals or Metalloids.

PROPERTIES.—Non-metals or metalloids possess no characteristics of metals, combine with oxygen into oxides (which when combined with water have acid properties), and in all other chemical and physical respects differ from metals.

Certain members of both groups show closely similar properties, and both metals and non-metals are divided into natural subgroups, whose members show a relation between atomic weights and properties—*e. g.*, the metals lithium, sodium, and potassium resemble each other in many *properties* and form a natural group. The sum of the atomic weights

of lithium and potassium ($7 + 39 = 46$) divided by 2 gives 23, the atomic weight of sodium.

Mendelejeff's Periodic Law.

This expresses the relation between atomic weights and properties of elements. "*When the elements are arranged in the order of their atomic weights, elements having similar chemical and physical properties will be found at regular intervals of the series.*" Mendelejeff's table shows elements arranged in vertical columns in the order of atomic weights, forming eight groups and twelve series, and those of similar physical and chemical properties fall into the same group.

The ALKALI METALS, are in Group I.

The HALOGENS, are in Group VIII.

AFFINITY FOR OXYGEN increases periodically from *left to right*, and AFFINITY for HYDROGEN increases from *right to left—e. g.*, *sodium* of Group I. forms Na_2O; *aluminum*, of Group III., Al_2O_3; *phosphorus*, of Group IV., P_2O_5; *chlorine*, of Group VII., HCl; *sulphur*, of Group VI., H_2S; and *silicon*, of Group IV., H_4Si.

SMALL PERIODS.—On account of this periodic change in properties of the elements of the same series, each subseries is called a *small period*.

BASIC PROPERTIES.—Oxides in Group I. are strong bases; in Group II. not so strong; in Group III. still weaker; in Group IV. indifferent or weakly acid.

ACID PROPERTIES then become marked in Groups V., VI., and VII., serially.

INTERSERIAL RELATIONS.—In the same group, members of the *even* periods resemble each other more closely than they do the members of the *uneven* periods, and *vice versa—e. g.*, in Group II. *calcium*, *strontium*, and *barium*, of the periods 4, 6, and 8, resemble one another more closely than they resemble *magnesium*, *zinc*, and *cadmium*, of the series 3, 5, and 7. *For this reason the elements in each group are arranged in two columns so that the elements of similar properties stand together.*

INTERMEDIATE SERIES.—In arranging the elements in this manner certain elements were found not to fit into *any* of the *groups* and *series* properly, but to form a series of three elements. whose properties made them fall as an intermediate series following the even periods 4, 6, and 10.

LARGE PERIODS are an *uneven*, an *even*, and an *intermediate* series taken together, and when *complete* comprise 17 elements.

BLANK SPACES.—These in many groups are apparently objectionable features of the table, but Mendelejeff stated that they were left for elements yet undiscovered. Since that time three elements, *scandium*, *gallium*, and *germanium*, have been discovered. Their general properties and atomic weights assigned them to their present places, previously blank. Argon and helion, recently discovered, have not been sufficiently investigated to give them a place in the table.

Mendelejeff's Periodic System.[1]

Series.	Group I.		Group II.		Group III.		Group IV.		Group V.		Group VI.		Group VII.		Group VIII.
	R_2O		RO		R_2O_3		RH_4 RO_2		RH_3 R_2O_5		RH_2 RO_3		RH R_2O_7		RO_4
1		H, 1													
2	Li, 7		Be, 9		B, 11		C, 12		N, 14		O, 16		F, 19		
3		Na, 23		Mg, 24		Al, 27		Si, 28		P, 31		S, 32		Cl, 35	Fe, 56; Ni, 59; Co, 59
4	K, 39		Ca, 40		Sc, 44		Ti, 48		V, 51		Cr, 52		Mn, 55		
5		Cu, 63		Zn, 65		Ga, 70		Ge, 72		As, 75		Se, 79		Br, 80	
6	Rb, 85		Sr, 87		Y, 89		Zr, 90,		Cb, 94		Mo, 96				Ru, 101; Rh, 103; Pd, 106
7		Ag, 108		Cd, 112		In, 113		Sn, 119		Sb, 120		Te, 125		I, 127	
8	Cs, 133		Ba, 137		La, 139		Cl, 142		Di, 145		Tl, 125				
9										Er, 166					
10					Yb, 173				Ta, 182		W, 184				Os, 190; Ir, 192; Pt, 194
11		Au, 196		Hg, 200		Tl, 204		Pb, 207		Bi, 208					
12							Th, 237				U, 239				

[1] In this table the decimals are omitted in the atomic weights.

Physical Classification.

The physical properties of elements vary: most elements are solid at ordinary temperatures; *mercury* and *bromine*, liquids; and five of the more important elements, gases (*hydrogen*, *oxygen*, *nitrogen*, *chlorine*, and *fluorine*).

Probably all, certainly most, of the solid elements form crystals under specific conditions; some are polymorphous, and others both amorphous and crystallizable.

Allotropy.

This is the property of elements to exist in different conditions.

ALLOTROPIC MODIFICATIONS are the different forms assumed—*e. g.*, carbon exists, *crystallized* as diamond and graphite, and *amorphous* as charcoal.

Allotropy is also possessed by some gases—*e. g.*, oxygen exists in *two* forms, *common* oxygen and *ozone*, differing physically and chemically. Allotropy is due to differences in the arrangement of the molecules in each form—*e. g.*, ozone has *three* atoms of oxygen in each molecule, and ordinary oxygen *two*.

A few elements have odor and taste; several, distinct colors; but most, no odor, taste, and no color.

Electrical Classification.

The electrical conditions of the different elements have been determined by electrolysis. Those collected at the positive pole are said to be NEGATIVE, and those at the negative pole, POSITIVE.

As a rule, the metals are electropositive, and non-metals electronegative. In the table below, each element is negative to those following and positive to those preceding.

Table of Elements Electrically Classified.

Negative End.

Oxygen.	Molybdenum.	Osmium.
Sulphur.	Tungsten.	Iridium.
Nitrogen.	Boron.	Platinum.
Fluorine.	Carbon.	Rhodium.
Chlorine.	Antimony.	Ruthenium.
Bromine.	Tellurium.	Palladium.
Iodine.	Tantalum.	Mercury.
Selenium.	Columbium.	Silver.
Phosphorus.	Titanium.	Copper.
Arsenic.	Silicon.	Uranium.
Chromium.	Hydrogen.	Bismuth.
Vanadium.	Gold.	Tin.

Indium.	Didymium.	Strontium.
Lead.	Cerium.	Barium.
Cadmium.	Thorium.	Lithium.
Thallium.	Zirconium.	Sodium.
Cobalt.	Aluminum.	Potassium.
Nickel.	Erbium.	Rubidium.
Iron.	Yttrium.	Cæsium.
Zinc.	Glucinum.	
Manganese.	Magnesium.	**Positive End.**
Lanthanum.	Calcium.	

The Chemical Characters of Elements.

The Nascent State of Elements.

This is the condition of elements at the instant their atoms leave molecules, and before the atoms have had time to combine with other atoms. During it they combine with greater energy than after having united to form molecules—*e. g.*, if hydrogen gas be collected and passed through a solution of arsenous oxide (As_2O_3) no reaction occurs; if, however, it be evolved *in a solution of arsenous oxid*, the *nascent* hydrogen atoms have sufficient energy to combine with both As and O atoms and form arsenuretted hydrogen (AsH_3) and water.

The Loss of Elementary Properties in Compounds.

CHEMICAL COMPOUNDS.

In combining into compounds, elements generally lose their characteristic properties. No general rule indicating the properties of the compounds can be given because of their great number and the varieties of their properties: usually a gaseous compound contains one gaseous element, and the union of two solid elements generally results in a solid, rarely a liquid, and never a gas at ordinary temperature.

Classification of Compounds.

No perfect system of classification of compounds has ever been devised. A few groups or classes of compounds exist, whose properties are so distinct and characteristic that a substance belonging to any one of them may be easily recognized. These groups are the *acids*, *bases*, and *neutral salts*.

Acids.

DEFINITION.—ACIDS are substances formed by the combination of *hydrogen* with an electronegative element or group of elements.

PROPERTIES.—*1.* When soluble in water they have a sour taste. 2.

They change the color of several organic substances, litmus being changed from blue to red ("*acid reaction*"). 3. The hydrogen is replaceable by metals. The resulting compound is called a *salt;* and according to the number of hydrogen atoms that may be replaced by metals acids are *monobasic*, *dibasic*, and *tribasic*. *Monobasic acids form normal salts only.*

Bases.

DEFINITION.—Bases are chemically the opposite of acids, are generally compounds of electropositive elements with oxygen, or oftener oxygen and hydrogen—*e. g.*, CaO, calcium oxide, and NaOH, sodium hydroxide.

PROPERTIES.—1. With acids the bases form salts—*e. g.*, sodium hydroxide and hydrochloric acid combine chemically into water and the salt sodium chloride: $NaOH + HCl = H_2O + NaCl$.

2. Bases soluble in water restore the color of organic substances changed by acids—*e. g.*, litmus is restored from red to blue ("*alkaline reactions*").

3. When soluble in water they taste of lye.

According to the number of hydrogen atoms that a base can replace in an acid, it is termed *monacid*, *diacid*, *triacid*, etc.

Salts.

DEFINITION.—Salts are formed when the hydrogen of an acid is replaced by a metal or a basic radical.

Normal salts have all the replaceable hydrogen of an acid replaced—*e. g.*, sodium chloride, NaCl, and sodium sulphate, Na_2SO_4, generally are *neutral* to litmus, but may be either *acid* or *alkaline*. Potassium carbonate, K_2CO_3, is *alkaline*, cupric sulphate, $CuSO_4$, *acid*, respectively, because carbonic acid is very weak and unable to overcome the alkalinity of the strong potassium hydroxide, and sulphuric acid is very powerful and outbalances the weak base cupric hydroxide.

Acid salts have only a portion of the replaceable hydrogen of acids replaced—*e. g.*, $NaHSO_4$, Na_2HPO_4, and $CaHPO_4$, are *generally acid* to litmus, but *frequently neutral*, and *occasionally alkaline*—*e. g.*, ordinary sodium phosphate, Na_2HPO_4, is slightly alkaline to litmus.

Basic salts (oxy-salts or sub-salts) contain an excess of the base over the quantity required for the formation of a neutral salt, and are considered as derived from bases by replacement of *part* of their hydrogen by acid radicals—*e. g.*, bismuth hydroxide has one, two, or three replaceable hydrogen atoms: if all three are replaced, as in bismuth nitrate, $Bi(NO_3)_3$, a normal salt results; if only one or two, as in $Bi{<}{NO_3 \atop (OH)_2}$ or $Bi{<}{(NO_3)_2 \atop OH}$, basic salts occur.

Double salts are formed by replacing the hydrogen of an acid by two metals, as potassium-sodium sulphate, $KNaSO_4$.

Residues, Radicals or Compound Radicals.

DEFINITION.—These are *unsaturated* groups of atoms, obtained by removal of one or more atoms from saturated groups, capable of combining actively with elements or other radicals, but having no separate existence—*e. g.*, the residue hydroxyl (OH) enters into many compounds, as sodium hydroxide (NaOH); calcium hydroxide, $Ca(OH)_2$, etc.

The radical (OH) is formed by the removal of one atom of hydrogen from the saturated compound H_2O, making one bond unsaturated, and the radical univalent. According to the number of bonds unsaturated, *bivalent*, *trivalent*, etc., radicals exist.

Neutral Substances.

DEFINITION.—These show neither acid nor basic properties, but many are known possessing limited characteristics of both acids and bases.

The Process of Neutralization.

DEFINITION.—When an acid and a base are mixed together so that the characteristic properties of each disappear they are said to have been neutralized.

TO EXPLAIN NEUTRALIZATION, it is necessary to consider the condition of the atoms in the molecules of acids and bases. Hydrogen replaceable by metals is an essential constituent of all acids, but the property of being replaced by metals is not a constant function of hydrogen atoms, as is shown by the fact that while iron can liberate hydrogen from sulphuric acid, the metal has no effect on oil of turpentine ($C_{10}H_{16}$). The hydrogen of acids is therefore in some peculiar condition which imparts to acids their characteristics, because when the hydrogen of an acid is replaced all its special features disappear. Each acid has its own properties, imparted by the element or group of elements combined with hydrogen.

This is proved by the fact that any acid will show certain definite reactions whether free or neutralized—*e. g.*, silver nitrate forms silver chloride when added to hydrochloric acid, and also when added to sodium chloride; barium chloride forms barium sulphate when added to free or neutralized sulphuric acid.

The characteristic action of an acid is supposed to be due to the separation (*dissociation*) of its constituent ions, when dissolved in water, *the hydrogen being the anion and the acid residue the cation* (see page 55). The cation may be a single atom or a group of atoms—*e. g.*, hydrochloric acid is dissociated into hydrogen and chlorine, and sulphuric acid into hydrogen and the group SO_4:

$$HCl = \underset{\textit{Positive.}}{H} + \underset{\textit{Negative.}}{Cl};$$

$$H_2SO_4 = \underset{\textit{Positive.}}{H} + \underset{\textit{Negative.}}{SO_4}.$$

The radical hydroxyl (OH) is characteristic of all soluble metallic bases, and when dissolved in water these bases dissociate into anions of the metals and cations of hydroxyl—*e. g.*, sodium hydroxide dissociates thus,

$$\text{NaOH} \quad = \quad \overset{\textit{Positive.}}{\text{Na}} \quad + \quad \overset{\textit{Negative.}}{\text{OH}}.$$

The dissociation of salts is similar to the above; the metal or basic radical forms the anion and the acid radical the cation—*e. g.*,

$$\text{NaCl} \quad = \quad \overset{\textit{Positive.}}{\text{Na}} \quad + \quad \overset{\textit{Negative.}}{\text{Cl}}.$$

In these dissociated ions the elements are certainly not in their ordinary condition, but the difference between these conditions is in essence not known.

Neutralization depends on the removal of the characteristic hydrogen from acids, and of the characteristic hydroxyl from alkalies. These constituents then combine to form water, which does not dissociate.

The neutralization of hydrochloric acid by sodium hydroxide is shown by the formula:

$$\underbrace{\overset{+}{\text{H}} + \overset{-}{\text{Cl}}} + \underbrace{\overset{+}{\text{Na}} + \overset{-}{\text{OH}}} = \underbrace{\overset{+}{\text{Na}} + \overset{-}{\text{Cl}}} + \underbrace{\overset{+}{\text{H}} + \overset{-}{\text{HO}}} = \text{NaCl} + \text{H}_2\text{O}.$$

The second reaction in this equation represents the condition assumed to exist in the substances formed during it as long as they remain in solution. When the solution is concentrated by evaporation, the ions of Na and Cl combine into molecules of NaCl, common salt.

QUESTIONS.

What is the total number of elements known, and how many are of great and general importance?

What are the characteristics of metals and metalloids?

What is meant by a natural group of elements?

What is Mendelejeff's periodic law?

What constitutes a period or a series in this law?

What constitutes a large period?

What is allotropy?

Give examples of allotropic modification.

What is meant by the nascent state?

What are the characteristics of acid, basic, and neutral substances?

What are salts?

What are meant by normal, basic, and acid salts?

What are meant by monobasic, dibasic, and tribasic acids?

What is a residue or radical?

Explain the process of neutralization.

What is the chemical basis of neutralization?

What are double salts? Give examples.

What are uni-, bi-, and trivalent radicals? Give examples.

What are mono-, di-, and triacid bases? Give examples.

CHAPTER III.

CHEMICAL COMBINATION AND DECOMPOSITION.

DEFINITIONS.—Chemical change, like all other natural phenomena, is due to motion, which may be produced by the *physical* action of heat, light, or electricity, or by the *chemical* action of one substance upon another, whereby the atoms better satisfy their affinities by rearranging themselves into new substances.

Chemical reaction is therefore the mutual action of two substances producing chemical change. The term is usually limited to the study of the substances decomposed or formed (*decomposition* and *composition*).

Reagents are substances used to bring about chemical reactions.

Combination into Compounds under Heat.

By exciting increased molecular vibration, heat causes chemical combination, the molecules being brought more closely under each other's attraction.

Decomposition of Compounds under Heat.

The breaking of mercuric oxide into mercury and oxygen by heat has been mentioned; similarly heat decomposes other compounds into elements or into simpler combinations—*e. g.*, heat converts cupric hydroxide into cupric oxide and water:

$$Cu(OH)_2 + \text{Heat} = CuO + H_2O,$$

in which one molecule of five atoms is split into two molecules of two and three atoms, respectively.

Heat causes chemical decomposition by increasing molecular vibration, which very likely makes the molecule less stable and weakens the atomic attraction. The temperature of decomposition varies with the nature of the substances.

Chemical Combination and Decomposition by Electricity and by Light.

Examples of these reactions will be considered (pages 44, 55, 56).

Combination and Decomposition by the Mutual Reaction of Substances.

For the exchange of atoms between two substances in chemical reaction, *free molecular* motion is required, consequently *direct chemical reac-*

tion between solids is rare. When solids of high chemical affinity are very finely powdered and well mixed, chemical combination may result. Usually one or both the solids are dissolved in a liquid, or are converted into a liquid or a gas by some means.

Substances may act on each other in three ways:

(1) DIRECT COMBINATION—*e. g.*, $H + Cl = HCl$;

(2) DIRECT OR SINGLE DECOMPOSITION—*e. g.*, $HCl + K = KCl + H$;

(3) INDIRECT OR DOUBLE DECOMPOSITION—*e. g.*, $KCl + AgNO_3 = AgCl + KNO_3$.

The formulas given in (1) and (2) are not entirely correct, because they represent the action as between free *atoms*, not between *molecules*. Therefore the first would be correctly written:

$$HH + ClCl = 2HCl;$$

and the second:

$$2HCl + KK = 2KCl.$$

CHEMICAL BASIS.—*The atoms of a certain element have greater affinity for those of one element than of another—e. g., the affinity between chlorine and potassium is greater than that between chlorine and hydrogen.*

No rule governing this affinity can be devised, because of variations in affinity shown by the same elements under different conditions —*e. g.*, steam passing over red-hot iron will form oxide of iron and hydrogen:

$$3Fe + 4H_2O = Fe_3O_4 + 8H.$$

This indicates that iron and oxygen have greater affinity for each other than hydrogen and oxygen when combined as steam.

When hydrogen is passed over hot oxide of iron, water and metallic iron are formed:

$$Fe_3O_4 + 8H = 4H_2O + 3Fe.$$

This, the exact reverse of the former reaction, indicates an opposite condition of affinities.

The Law of Mass-action.

The alteration of the affinity of oxygen, that is indicated by the last two reactions, is explained by the law of mass-action.

"*Chemical action is proportional to the active mass of each substance taking part in the change.*" When a great excess of iron is acted on by steam the iron decomposes the water, and when oxide of iron is acted on by an excess of hydrogen iron and water are formed.

The Laws of Berthollet.

These govern certain chemical decompositions, and are:

I. When two or more substances are together in solution, which by rearrangement of their atoms can form a substance insoluble in the solvent, this insoluble resultant substance will form and separate as a precipitate—e. g., the formation of insoluble barium sulphate by mixing solutions of barium chloride and sodium sulphate.

II. When two substances are brought together, which by rearrangement of their atoms can form a substance gaseous at the temperature of the experiment, this gaseous resultant substance will form and escape as a vapor—e. g., the liberation of ammonia when ammonium chloride and sodium hydroxide are mixed.

ANALYSIS AND SYNTHESIS.

Chemical reactions may be analytical or synthetical.

DEFINITIONS.—*Analytical* reactions decompose compounds into their elements or simpler compounds.

Analysis is investigation by this method.

Synthetical reactions combine compounds into more complex substances.

Synthesis is investigation by this method.

CHEMICAL EQUATIONS AND STOICHIOMETRY.

DEFINITION.—*Chemical formulas* have been mentioned, the *molecular* on page 66, the *graphic* on page 69.

Chemical Equations.—The molecular formulas, when combined into equations, *mathematically* represent chemical reactions.

Stoichiometry is that part of chemistry dealing with these equations.

FORMATION OF CHEMICAL EQUATIONS.—In their *left-hand* member the formulas of substances brought together are written, connected with a *plus* (+) sign, and the results of the reactions, in their *right-hand* member—*e. g.*,

$$CaCO_3 + H_2SO_4 = CaSO_4 + H_2O + CO_2.$$

This equation indicates that when a molecule of calcium carbonate and of sulphuric acid are brought together the reaction forms one molecule each of calcium sulphate, water, and carbon dioxide.

If two or more molecules enter the reaction, their number is indicated by the appropriate figure placed before the formulas, or symbols—*e. g.*,

$$CaCO_3 + 2HCl = CaCl_2 + H_2O + CO_2.$$

This expression indicates the reaction between *one* molecule of calcium carbonate and *two* of hydrochloric acid; and

$$3BaCl_2 + 2H_3PO_4 = Ba_3(PO_4)_2 + 6HCl$$

indicates the reaction between *three* molecules of barium chloride and *two* of phosphoric acid.

PROOF OF CHEMICAL EQUATIONS.—*As matter is indestructible, it is obvious that the sum of the weights of the molecules in one member of an equation must equal that of the other, and every correct chemical equation is therefore correct mathematically.*

Practically these equations are of great value, as by their use the relative proportions are calculated in which substances must be mixed to produce certain other substances—*e. g.*, the first foregoing equation indicates that a mixture of calcium carbonate 100 parts and sulphuric acid 98 parts will form 136 parts, by weight, of calcium sulphate, 18 parts of water, and 44 parts of carbon dioxide (the molecular weight of each):

$CaCO_3$ +	H_2SO_4 =	$CaSO_4$ +	H_2O +	CO_2
Ca = 40	2H = 2	Ca = 40	2H = 2	C = 12
C = 12	S = 32	S = 32	O = 16	2O = 32
3O = 48	4O = 64	4O = 64	18	44
100	98	136		

Equations may also indicate the actual weights of substances required to produce a definite weight of a product—*e. g.*, to produce 132 lb. of carbon dioxide, 300 lb. of calcium carbonate and 294 lb. of sulphuric acid will be required:

$$CaCO_3 + H_2SO_4 = CaSO_4 + H_2O + CO_2$$
$$100 + 98 = 136 + 18 + 44$$

$$100 : 44 :: x : 132 \text{ lb.}$$
$$44x = 13200$$
$$x = 300 \text{ lb.}$$

$$98 : 44 :: x : 132 \text{ lb.}$$
$$44x = 12936$$
$$x = 294 \text{ lb.}$$

The third and fourth terms of any such equation will always be of the same denomination. In writing chemical equations it is necessary to remember that equivalent quantities replace one another—e. g., two atoms of a *univalent* element replace *one* atom of a *bivalent* element, *three univalent* replace *one trivalent*, *three bivalent* replace *two trivalent*, etc.

QUESTIONS.

How may chemical decomposition be produced?
What is a chemical reaction?
Mention three forms of reaction.
What is a reagent?
Give an example of mass-action.
State the laws of Berthollet.

What is the difference between analysis and synthesis?
What facts must be observed in writing equations?
Why are equations correct mathematically when correct chemically?
Complete the following equations, indicating the correct number of molecules of the factors in the first six:

$H_2'SO_4 + Na' =$
$Ba''Cl_2 + Ag'NO_3 =$
$Al'''Cl_3 + Ag'NO_3 =$
$Bi'''(NO_3)_3 + Na'OH =$
$Fe_2'''Cl_6 + K'OH =$
$Fe_2'''(SO_4)_3 + Ba'''Cl_2 =$
$Na_2'CO_3 + Ca''(OH)_2 =$
$K'Cl + Ag'NO_3 =$
$Cu''SO_4 + K'OH =$
$Pb''(NO_3)_3 + Na_2'CO_3 =$
$Mg''SO_4 + Ba''Cl_2 =$
$Zn''SO_4 + Na'OH =$

CHAPTER IV.

CHEMICAL NOMENCLATURE.

General Purpose.

IN modern chemical nomenclature, the name of a compound expresses our view of its constitution or properties.

The Significance of the Suffix -Ide.

Compounds, formed by the union of two elements uniting in one proportion only, are named by adding the suffix *ide* to the name of the *negative* atom—*e. g.*,

Sodium chloride and potassium iodide, for the compounds NaCl and KI.

The Significance of the Prefixes Mono-, Di-, etc.

When several compounds are formed by the union of the *same* elements in *different* proportions, the prefixes *mono-*, *di-*, *tri-*, *tetra-*, etc., are joined to the name of the *negative* atom—*e. g.*,

Carbon monoxide, CO; carbon dioxide, CO_2; sulphur dioxide, SO_2; phosphorus trichloride, PCl_3; and phosphorus pentachloride, PCl_5.

The Significance of the Suffixes -Ic and -Ous.

The different substances formed by the union of the *same* elements *in two proportions* are named by adding the suffix *ic* to the name of the

positive element for the name of the compound in which this element has the *greater* valence, and *ous*, of the compound in which it has the *lower* valence—*e. g.*,

Phosphoric oxide for P_2O_5, and phosphorous oxide for P_2O_3; ferric oxide for Fe_2O_3; and ferrous oxide for FeO.

When *ous* and *ic* oxides of the *same* element form acids by combining with water, the acids are distinguished by the same suffixes—*e. g.*,

$$\underset{\text{Sulphurous oxide}}{SO_2} + H_2O = \underset{\text{— Sulphurous acid}}{H_2SO_3}$$

$$\underset{\text{Sulphuric oxide}}{SO_3} + H_2O = \underset{\text{— Sulphuric acid}}{H_2SO_4}$$

The Significance of the Prefixes Mono-, Sesqui-, and Sub-.

The prefixes *mono-* and *sesqui-* are sometimes joined to the name of the *positive* element, signifying the same as the suffixes *ous* and *ic*—*e. g.*,

Monoxide of iron, FeO, and sesquioxide of iron, Fe_2O_3.

The prefix *sub* is sometimes joined to the *negative* element to indicate the compound in which the *positive* element shows the *lower* valence—*e. g.*,

Suboxide of copper, for cuprous oxide, Cu_2O.

Nomenclature of Compounds in Four Proportions.

When the *same* elements unite in *four* proportions the compounds are named by joining the prefixes *per* and *hypo*, to the name of the *positive* element for the name of *most complex* and *simplest* compounds, respectively, and the suffixes *ic* and *ous* for the intermediate ones in the order of complexity—*e. g.*,

Hypochlorous oxide for Cl_2O, phosphorous oxide for P_2O_3, nitric oxide for N_2O_5, and perchloric oxide for Cl_2O_7.

The *acids* formed by combinations of *hypo* and *per* oxides with water are designated by the same suffixes—*e. g.*,

$Cl_2O + H_2O = 2HClO$, hypochlorous acid,

$Cl_2O_7 + H_2O = 2HClO_4$, perchloric acid.

The *salts* of these acids are named by adding to the name of the *leading negative* atom or radical the prefixes *hypo*, and *per* in the case of *hypo* and *per* acids, respectively, and the suffixes *ite* and *ate* in the case of *ous* and *ic* acids, respectively—*e. g.*:

$HClO + NaOH = H_2O + NaClO$, sodium hypochlorite;

$HClO_2 + NaOH = H_2O + NaClO_2$, sodium chlorite;

$HClO_3 + NaOH = H_2O + NaClO_3$. sodium chlorate;
$HClO_4 + NaOH = H_2O + NaClO_4$, sodium perchlorate.

The Significance of the Prefixes Proto- and Per-.

Occasionally, particularly in older works on chemistry, the suffixes *-ous* and *-ic*, *-ite* and *-ate*, of the positive elements are replaced by the prefixes *proto-* and *per-*, joined to the negative element—*e. g.*,

Protochloride of iron for ferrous chloride ($FeCl_2$) and perchloride of iron for ferric chloride (Fe_2Cl_6).

The Significance of the Term Anhydride.

Anhydride applies to those oxides which unite with water to form acids—*e. g.*,

Sulphurous anhydride for SO_2, phosphoric anhydride for P_2O_5. etc.

QUESTIONS.

What is the significance of the suffix -ide in chemical nomenclature?
What is meant by the prefixes mono-, di-, tri-, tetra-, penta-?
What is the significance of the syllables -ous and -ic, -ite and -ate?
What is meant by the syllables per- and hypo-, proto- and per-?
What is an anhydride?
Distinguish -ous from -ic acids.
What is the difference between mono-, sesqui-, and sub- compounds?
Discuss the names of compounds from elements uniting in four proportions, indicated by hypo- and per-, -ic and -ous.
What are hypo- and per- acids?
What are hypo- and per-, -ite and -ate salts?

CHAPTER V.

HYDROGEN AND COMPOUNDS OF HYDROGEN.

HYDROGEN (ὕδορ, water; γεννάω, to generate).

S.,[1] H; **V.**,[1] 1; **A. W.**,[1] 1.

HYDROGEN is the standard of atomic weight and valence. **A liter of it at standard temperature and pressure (see page 28), weighs 0.0899 gram, which is the unit of weight for gases, called a crith.**

[1] For simplicity the following abbreviations will be used in connection with the description of elements and compounds: S. = symbol, V. = valence, A. W. = atomic weight, M. F. = molecular formula, G. F. = graphic formula, M. W. = molecular weight.

HISTORY.—Its name was given by Lavoisier, because its combustion forms water.

OCCURRENCE.—*Terrestrially* hydrogen occurs chiefly as a constituent of water, of *all* acids, of most animal and vegetable compounds, and, in small quantity, of the mixture of gases produced by organic decomposition, as in volcanic gases, natural gas, and the gases of the animal intestine, and *celestially*, of the sun and certain stars.

PREPARATION.—1. By electrolysis of water (see page 55); 2, by decomposition of water by metals; 3, of acids by metals; 4, of strong solutions of sodium or potassium hydroxide by zinc or aluminum.

Water is decomposed by certain metals—*e. g.*, sodium or potassium, at ordinary temperatures, and by other metals, like iron, at red heat.

Zinc or iron and dilute hydrochloric or sulphuric acid are generally used to generate hydrogen.

These reactions are respectively represented by the following equations:

$$1.\quad H_2O = 2H + O \ (\textit{Electrolysis}),$$

$$2.\ \begin{cases} K + H_2O = KOH + H, \\ 3Fe + 4H_2O = Fe_3O_4 + 8H, \end{cases}$$

$$3.\ \begin{cases} Zn + H_2SO_4 = ZnSO_4 + 2H, \\ Fe + 2HCl = FeCl_2 + 2H, \end{cases}$$

$$4.\ \begin{cases} Zn + 2KOH = K_2ZnO_2 + 2H, \\ Al + 3NaOH = Na_3AlO_3 + 3H. \end{cases}$$

PHYSICAL PROPERTIES.—A tasteless, odorless, and colorless gas; the lightest of all known substances; at a temperature of —253° C. a colorless liquid, and at about —256° C. (calculated) a solid.

CHEMICAL PROPERTIES.—Hydrogen has great affinity for oxygen, removing it from many oxides. When passed over red-hot oxide of iron or copper it combines with the oxygen into water and frees the metals—*e. g.*,

$$CuO + 2H = Cu + H_2O.$$

DEOXIDATION or REDUCTION is the term applied to this process.

DEOXIDIZING or REDUCING AGENTS have the power of removing oxygen from certain of its compounds.

Hydrogen burns in air or oxygen with a non-luminous colorless or pale-blue flame, whose heat is greater than that of the combustion of an *equal weight* of any other substance.

A mixture of hydrogen and oxygen, or air, when ignited, explodes, forming water, H_2O. All air should therefore be expelled from the hydrogen generator before igniting the gas.

THE COMPOUNDS OF HYDROGEN.

The compounds of hydrogen with other elements will be considered under these elements to avoid repetition.

QUESTIONS.

Mention processes for obtaining pure hydrogen.

Show by symbols the actions of sodium on water; iron on sulphuric acid; zinc on potassium hydroxide.

What are the atomic weight and valence of hydrogen?

What is a crith?

How is hydrogen found in nature?

What are the physical and chemical properties of hydrogen?

How many grams of zinc and sulphuric acid are required to form 50 grams of hydrogen?

CHAPTER VI.

OXYGEN, ALLOTROPIC FORMS OF OXYGEN, AND COMPOUNDS OF OXYGEN.

OXYGEN (ὀξύς, acid; γεννάω, to generate).

S., O; **V.**, 2; **A. W.**, 16 (15.96).

History.—Discovered independently by Scheele, of Sweden, in 1773, and by Priestley, of England, in 1774.

The name was devised by Lavoisier, who first recognized its nature.

Occurrence.—*Terrestrially*, in greater quantity than any other element; it forms probably 45 per cent., certainly 33 per cent. of the total weight of the earth.

Table of the Quantitative Terrestrial Occurrence of Oxygen.

The atmosphere = 20 per cent. *free* oxygen.

Water = $88\frac{8}{9}$ per cent. *combined* oxygen.

$\left.\begin{array}{l}\text{Rocks and minerals}\\ \text{Constituents of earth}\end{array}\right\}$ = 30 to 50 per cent. *combined* oxygen.

Animal and vegetable life = varying quantities *combined* oxygen.

Preparation.—1. By decomposing oxides, as of gold, silver, platinum, or mercury by heat (see page 106):

$$HgO = Hg + O,$$
$$Ag_2O = 2Ag + O.$$

2. By heating potassium chlorate, forming potassium chloride and oxygen:

$$2KClO_3 = KClO_4 + KCl + 2O.$$

This decomposition occurs at comparatively low temperature; after its completion, with considerable rise of temperature the perchlorate *decomposes*:

$$KClO_4 = KCl + 4O$$

If potassium chlorate be mixed with from 30 to 50 per cent. of manganese dioxide, its entire oxygen is obtained at lower temperature than by heating the chlorate alone. The quantity of manganese dioxide is unchanged at the end, apparently having taken no active part in the decomposition. No satisfactory explanation of this has ever been given.

3. By heating manganese dioxide to redness in an iron retort:

$$3MnO_2 = Mn_3O_4 + 2O;$$

4. By heating a concentrated solution of calcium hypochlorite with a small quantity of a cobaltous salt:

$$Ca(ClO)_2 = CaCl_2 + 2O.$$

The following are important among many other methods of obtaining oxygen:

5. By decomposing water by electricity;
6. By heating dichromates, nitrates, and barium dioxide;
7. By adding water to a mixture of potassium ferricyanide and barium dioxide, and by mixing potassium permanganate and hydrogen dioxide in the presence of dilute sulphuric acid.

PHYSICAL PROPERTIES.—This colorless, odorless, and tasteless gas becomes a pale-blue liquid at —140° C. and a pressure of 300 atmospheres, with a critical temperature of —120° C. and boiling-point under ordinary pressure of —180° C.; is soluble in water about 3 volumes in 100.

CHEMICAL PROPERTIES.—Oxygen shows great affinity for other elements combining *directly* with most metals and non-metals, *indirectly*, chiefly with chlorine, bismuth, iodine, gold, silver, and platinum, and *not at all* with fluorine only of the important elements.

OXIDATION is the act of combination of oxygen with other substances and OXIDES its product.

COMBUSTION is oxidation or any other chemical action, generating sufficient heat to emit light.

Slow combustion is oxidation not accompanied by light, as in the human body.

The intensity of heat in combustion depends on its rapidity, but the total quantity of heat is the same whether the combustion be rapid or slow.

Combustion requires that a combustible substance and a supporter of combustion (ordinarily oxygen) act chemically on each other. Under certain conditions oxygen may be the combustible substance—a reversal of its usual relations.

The quantity of heat in oxidation is exactly equal to that required to decompose the substances formed—*e. g.*, the amount of heat evolved by combining mercury and oxygen into red oxide of mercury will be required to decompose it into mercury and oxygen.

The heat of combustion is a conversion of chemical into heat-energy.

Substances having the power of uniting with others are capable of doing chemical work, and possess chemical energy.

The kindling-point is the temperature of commencing combustion, and varies with different substances, from ordinary temperature to white heat.

The presence of free oxygen is not essential for oxidation, as certain oxygen compounds (*oxidizing agents*) readily give up their oxygen to substances having a greater affinity for it—*e. g.*, nitric acid, chromic acid, potassium permanganate, and potassium chlorate.

ALLOTROPIC FORMS OF OXYGEN.

Ozone.

S., O_3; **M. W.**, 46; **G. F.**, $\begin{matrix} O \\ | \\ O \end{matrix}\!\!>\!O$.

PREPARATION.—1. By passing non-luminous electric discharges through oxygen or atmospheric air; 2, by exposing to the air phosphorus, partially covered with water; 3, during a number of other chemical decompositions.

PHYSICAL PROPERTIES.—Ozone differs from oxygen in a peculiar odor, and stronger oxidizing power.

CHEMICAL PROPERTIES.—Its presence is recognized by its odor and its action on a piece of paper wet with solutions of potassium iodide and starch-paste. It liberates the iodine, which forms a dark-blue compound with the starch. Oxygen does not decompose potassium iodide.

Three volumes of oxygen are condensed to two when converted into ozone. For this reason it is assumed that three molecules of oxygen form two of ozone, and that the ozone molecule contains three atoms, common oxygen, but two atoms in its molecule:

$$3[O = O] = 2\left[\begin{matrix} & O & \\ \diagup & & \diagdown \\ O & \text{——} & O \end{matrix}\right].$$

COMPOUNDS OF OXYGEN.

Water (Hydrogen Monoxide).

M. F., H_2O; **M. W.**, 18; **G. F.**, $\begin{matrix} H \\ H \end{matrix}\!\!>\!O$.

OCCURRENCE.—*Terrestrially*, this very important oxygen-compound is widely distributed, and in all of three states of aggregation, *gaseous* in the atmosphere, *liquid* in the oceans, rivers, and other bodies, *solid*, as ice, snow, and water of crystallization, *liquid* in animals and plants in high percentage by weight—*e. g.*, 70 per cent. of the human body, and *75 per cent. of potatoes.*

Classification of Water.

For the conveniences of description, water may be considered as:
1. Natural water, suitable for domestic and commercial purposes;
2. Potable water, adapted for drinking by man;
3. Mineral water, possessed of special medicinal value;
4. Distilled or pure water, suitable for chemical laboratory use; and—
5. Crystallization-water or water of crystallization, essential to the process of crystallization in many substances.

Natural Water.

PROPERTIES.—It is never absolutely pure, but always contains *gases* or *solids* in solution, but is purest as rain-water, *collected after the air has been washed by an immediately previous shower.*

Water, when partially frozen, leaves most of its impurities in the uncongealed portion, therefore melted ice furnishes a comparatively pure water.

The waters of springs, wells, and rivers, differ widely in their dissolved substances, taken up during their passage through atmosphere, soil, and rocks.

Gases of Water.—These are oxygen, nitrogen, and carbon dioxide, chiefly from the atmosphere. *One hundred volumes of water* dissolve *two* of nitrogen, *one* of oxygen, and *one* of carbon dioxide.

Solids of Water.—These are *usually* sodium chloride, calcium sulphate, calcium carbonate and bicarbonate, *occasionally* potassium and magnesium salts, and traces of iron salts and silica.

Hard water contains 150 mg. per liter or more of calcium or magnesium bicarbonate or sulphate, will not dissolve soap, and consequently is not useful for domestic purposes.

Temporary hardness, due to bicarbonates, is removed by boiling, which converts the bicarbonates into insoluble carbonates.

Permanent hardness is due to the presence of sulphates, and is removed by adding sodium carbonate in excess.

Potable Water.

PROPERTIES.—A good drinking-water should have no odor, color, nor disagreeable taste. An absolutely pure water is not a good drinking-water, nor is one containing too much foreign matter.

STANDARDS OF POTABILITY.—*Inorganic Contamination.*—From two to four parts of inorganic solids, principally sodium chloride and calcium carbonate, with 100 volumes of carbon dioxide in 10,000 parts of water, will be contained by a good drinking-water, although more than the above solids is allowable.

Organic Contamination.—Organic, especially animal matter, is objectionable in drinking-water. Living organisms (germs), frequently derived from decomposing animal matter, and generally causing disease, render drinking-water dangerous.

APPROXIMATE TEST FOR ORGANIC MATTER.—Add a small quantity of sulphuric acid and enough potassium permanganate to give the mixture a distinct red color. On boiling the solution becomes decolorized if organic matter is present.

PRACTICAL DISINFECTION AND PURIFICATION.—By boiling actively and continuously for one hour and filtering through sand, charcoal, spongy iron, or some similar material, an unwholesome water may be made potable. The boiling destroys the germs and filtering removes the impurities. The materials retained by the filter may become a source of pollution, if the same filter is used continuously. This may be avoided by heating the filter material to red-heat once a day.

CHEMICAL EXAMINATION OF WATER.—This is not always sufficient to determine its potability.

Micro-organisms cannot always be detected by chemical, but can by bacteriological examination.[1]

Certain substances in sufficient quantity are indications of pollution, and chemical analysis to determine their presence or absence usually precedes the bacteriological examination.

While no fixed standard of chemical purity for drinking-water has been decided, most authorities agree that the quantities mentioned below are the maxima of admixture allowable in a liter of water.

Total solids include those *suspended* and *dissolved.*

Suspended solids occur in turbid water, and their quantity is determined by passing a liter of water through a previously dried and weighed filter; after drying again, the weight of the residue and filter are determined, and the weight of the filter subtracted, leaving the weight of suspended solids.

Allowed Total per Liter.—No definite quantity may be stated. This element of potability depends upon the other factors also present, and must be decided accordingly.

Dissolved Solids.—Evaporate to dryness a liter of clear filtered water in a weighed platinum dish and dry the residue at 110° C. The weight of the residue is the quantity of dissolved solids. *Allowed total:* 500 mg. per liter.

Organic matter is determined approximately by boiling with a solution of potassium permanganate of known strength. *Allowed total:* 1 liter of water should not decolorize 10 mg. of pure potassium permanganate when acidified with 5 cc. of sulphuric acid and boiled for ten minutes.

The possibility of the water containing deoxidizing agents not organic must be considered in this test.

Ammonia, free and *albuminoid*, are both evidence of organic pollution.

Free ammonia is formed by the action of bacteria on nitrogenous organic matter, and is generally found in water as ammonium carbonate.

[1] For the detailed bacteriological examination of water consult the volume on *Microscopy and Bacteriology.*

Albuminoid ammonia is formed by the oxidation of nitrogenous organic matter in water by potassium permanganate.

Tests.—To determine the quantities of these substances present, 500 c.c. of water are mixed with 5 c.c. of saturated solution of sodium carbonate in a suitable flask, which is connected with a condenser, whose outlet enters a receiver so as to prevent loss of ammonia. About 300 c.c. are distilled. This distillate, containing all the free ammonia, is diluted to 500 c.c. with distilled water, and then the solution Nesslerized (see below).

The liquid remaining in the flask is mixed with 5 grams of potassium permanganate and 10 grams of potassium hydroxide dissolved in 50 c.c. of distilled water, and the mixture distilled until 200 c.c. have passed over. The distillate is diluted to 500 c.c. and Nesslerized. *In making these determinations it is necessary that no ammonia be present in the atmosphere in which the work is carried out.*

Nessler's Reagent.—Dissolve 5 grams of potassium iodide in hot water, and to this add a hot solution of 2.5 grams of mercuric chloride in 10 c.c. of water. To this turbid red mixture a solution of 16 grams of potassium hydroxide in 40 c.c. of distilled water is added, and then enough distilled water to make the entire volume 100 c.c. On cooling, some crystals of mercuric iodide deposit and the clear solution is decanted as needed. *This reagent gives a yellow to brown color with dilute solutions of ammonia, and the intensity of the color is a measure of the quantity of ammonia present.*

Standard Ammonia Solution.—This is made for comparison, containing 3.137 grams of ammonium chloride, corresponding to 1 mg. of NH_3 in a liter. Just before use, 5 c.c. of this solution are diluted to 100 c.c., therefore 1 c.c. of the dilution contains 0.05 m.g. NH_3. The test is made in cylinders of colorless glass of the same height and diameter (100×30 mm.) marked at 50 c.c. In the first cylinder 0.5 c.c. of the standard ammonium chloride solution is placed, 1 c.c., 1.5 c.c., 2 c.c.. in the second, third, and fourth cylinders, respectively, and all are filled to the mark with water, making the contents of the cylinders correspond to water containing 0.5, 1, 1.5, and 2 mg. of NH_3 per liter; 50 c.c. of the suspected water are placed in a fifth cylinder, and to each of the five cylinders 1 c.c. of Nessler's reagent is added and well mixed with the contained solution. *By comparing the shade of color in the fifth cylinder with that in the other cylinders, containing known quantities of ammonia, that in the suspected water is estimated—most satisfactorily, if the cylinders are on white paper.*

Allowed total of free and albuminoid ammonia: 0.05 mg. and 0.1 mg. per liter, respectively.

Nitric acid is shown by the tests with brucine, diphenylamine, and pyrogallic acid (see chapter on Nitrogen). *Allowed total:* 10 mg. of nitrogen pentoxide (N_2O_5) per liter.

Nitrous acid is determined by using a solution of 1 gram of metaphenylenediamin in 200 c.c. of 2.5 per cent. sulphuric acid solution, by a

process analogous to Nessler's test for ammonia. The solution gives with nitrites a yellow color, whose intensity serves for their quantitative estimation.

Standard Nitrate Solution.—This is made by dissolving 0.406 gram of silver nitrate and 0.225 gram of potassium chloride in hot water, cooling, and making up to 1 liter. After filtering, 100 c.c. of the solution are diluted to 1 liter, equivalent to a solution of 10 mg. of N_2O_3 per liter.

Test.—This is made by using definite dilutions of the standard nitrate solution in four cylinders, and adding 1 c.c. of the metaphenylenediamin solution to each 50 c.c. of liquid and comparing the shades of color produced with that obtained by treating the suspected water with the same proportion of the reagent.

Allowed total: 0.05 mg. of nitrogen trioxide (N_2O_3) per liter.

Sulphates are determined by adding barium chloride in acid solution to from 100 to 250 c.c. of water. The precipitate is dried and weighed and the quantity of acid calculated from its weight.

Allowed total: 50 to 80 mg. of sulphur trioxide (SO_3) per liter.

Chlorides may occur in considerable quantities from strata containing salt deposits, or from the sea-coast or salt lakes, without indicating pollution.

Allowed total: When such a source cannot be determined, more than 15 *mg.* of *chlorine per liter shows unpotability by contamination from cesspools, sewage, etc.*

Test.—The quantity of chlorine is determined by titration with $\frac{N}{10}$ $AgNO_3$ (see Volumetric Analysis), using potassium chromate as an indicator.

Phosphates.—*Soluble* phosphates in water are almost positive proof of pollution with urine.

Test.—Ammonium molybdate test (see chapter on Phosphorus).

Mineral Waters.

These are spring-waters containing enough of one or more substances to impart to them a peculiar taste, and usually medicinal properties.

CLASSIFICATION OF MINERAL WATERS, according to their predominating constituents:

Acid, containing free mineral acids;
Alkaline, containing alkali bicarbonates;
Alum, containing alum;
Chalybeate, containing much iron carbonate or sulphate;
Bitter, containing much magnesium salts;
Effervescent, containing much free carbonic acid;
Hepatic or *Sulphur*, containing hydrogen sulphide.

Distilled Water.

PROPERTIES.—Properly obtained this is pure water. **From 1000** *parts of ordinary* water 700 of pure water may be obtained by **distillation,**

the first 100 parts of distillate containing the gases dissolved in the water should not be used, the next 700 distilled being pure, and the 200 parts undistilled containing the dissolved solids.

PHYSICAL AND CHEMICAL PROPERTIES.—Water is an inodorous, tasteless, transparent liquid, in thick layers of blue tint. It is perfectly neutral, but combines with both acid and basic substances. *Hydroxides* is the name of compounds often formed by the direct union of oxides with water—*e. g.*,

$$CaO + H_2O = Ca(OH)_2.$$

Water is the common solvent. This power usually increases with its temperature, but there are many exceptions to this rule—*e. g.*, the *solvent power for gas increases with pressure and diminution of temperature.*

Water of Crystallization.

Water of crystallization is water taken up by many salts in crystallizing, combined without alteration of atomic arrangement, and easily expelled by heat, mostly at 100°–120° C. It has a great influence on crystalline form.

EFFLORESCENCE.—Some substances on exposure to air lose their water of crystallization, and then their crystalline form, falling to powder—*i. e.*, *effloresce.*

DELIQUESCENCE is the property of certain substances of attracting water from the air and gradually becoming liquid.

HYGROSCOPIC substances absorb water from the air and become damp.

Hydrogen Dioxide (Hydrogen Peroxide).

$$\textbf{M. F.},\ H_2O_2;\ \textbf{M. W.},\ 34;\ \textbf{G. F.},\ \begin{matrix} O-H \\ | \\ O-H \end{matrix}.$$

Hydrogen peroxide is another very important combination of hydrogen and oxygen.

PREPARATION.—As a dilute aqueous solution. 1. By the action of acids on certain metallic dioxides in the presence of water, barium or sodium dioxide, which yield a portion of their oxygen to the water:

$$BaO_2 + H_3PO_4 + H_2O = BaHPO_4 + H_2O + H_2O_2.$$

Considerable water is necessary for the decomposition.

2. As 50 per cent. solution, by evaporation of this dilute solution at a temperature not exceeding 60° C.;

3. Pure by distilling the 50 per cent. solution *in vacuo*, passing over at 85° C.

PHYSICAL PROPERTIES.—Pure hydrogen dioxide is a colorless, oily liquid, Sp. G. 1.45, soluble in water, alcohol, and ether.

CHEMICAL PROPERTIES.—Decomposes slowly at the ordinary temperature, rapidly at 100° C., forming water and oxygen. A dilute solution is more stable and may be boiled without decomposition, but is decomposed by exposure to light and by many inert substances in powder—*e. g.*, fine metallic powders. All ferments decompose the solution. The presence of traces of free acids or of such substances as boroglyceride makes the solution more stable.

TESTS.—1. It decomposes by the addition of silver oxide into oxygen, water, and metallic silver:

$$H_2O_2 + Ag_2O = 2O + 2Ag + H_2O.$$

2. Added to a solution of potassium iodide containing some starch paste it gives a blue color, by forming iodized starch, by iodine liberated from the potassium iodide:

$$H_2O_2 + 2KI = 2I + 2KOH.$$

3. It decolorizes acidified solutions of potassium permanganate with the liberation of oxygen:

$$5H_2O_2 + 2KMnO_4 + 3H_2SO_4 = 10O + K_2SO_4 + 2MnSO_4 + 8H_2O.$$

This reaction is used in the quantitative estimation of hydrogen dioxide by volumetric solution of potassium permanganate.

THERAPEUTIC PROPERTIES.—With pus it causes effervescence, due to the liberation of oxygen, which usually combines with the pus-elements, especially the bacteria. It is one of the most active antiseptics known, and damages the tissues to a relatively less degree than do the other common purely chemical antiseptics.

Glycozone is a solution of hydrogen dioxide in glycerin.

U. S. P. STANDARD SOLUTION.—This contains 3 per cent. by weight of pure dioxide, corresponding to about 10 volumes of available oxygen to 1 of solution, and, like the pure dioxide, has antiseptic, caustic, and bleaching properties, usually acting as an oxidizing agent.

QUESTIONS.

When and by whom was oxygen discovered? By whom was it so named, and why?

Mention four methods by which oxygen may be obtained.

Why is manganese dioxide added to potassium chlorate in making oxygen?

How much oxygen may be obtained from 61.2 grams of potassium chlorate?

State the physical and chemical properties of oxygen.

What is combustion?

What is meant by combustible substances, and by supporters of combustion?

What is the kindling-point?

What are oxides, oxidizing and reducing agents?

What is ozone, and does it differ from oxygen? How is it formed?

What is the composition of water in parts by weight and by volume?

How may pure water be obtained?

What is hard water?
What are the causes of temporary and permanent hardness?
How may hardness of water be overcome?
What are the characteristics of good drinking-water?
How may pollution of water be detected by chemical analysis?
What are the characteristic constituents of the chief classes of mineral waters?
What are the physical and chemical properties of water?
What is water of crystallization?
What is meant by efflorescence, deliquescence, and hygroscopicity?
Show by equation the formation of hydrogen dioxide from barium dioxide.
State the chemical and physical properties of hydrogen dioxide. What is the strength of its U. S. P. solution?
Give three reactions for hydrogen dioxide. What is the nature of its clinical reaction? Why is it of very great clinical value?

CHAPTER VII.

NITROGEN AND COMPOUNDS OF NITROGEN.

S., N; **V.**, 3 or 5; **A. W.**, 14.

History.—Discovered by Rutherford in 1712, and named from the Greek *νιτρον*, nitre, and *γεννάω*, to produce.

Occurrence.—In nature, free in the atmosphere and combined as nitrates, ammonia, and in many organic substances.

Preparation.—1. In an *impure* condition, by burning phosphorus in air, confined in a bell-jar held over water: the phosphoric oxide formed is dissolved by the water, leaving the nitrogen in the jar.

2. *Pure* nitrogen, by heating ammonium nitrate in a glass retort:

$$NH_4NO_2 + \text{Heat} = 2H_2O + 2N.$$

Physical Properties.—Nitrogen is a colorless, tasteless, inodorous gas, slightly soluble in water.

Chemical Properties.—It is neither combustible nor a supporter of combustion; has very little affinity for other elements, does not enter into direct combination with scarcely any other substance; not being a supporter of combustion can not sustain animal life, but is not poisonous in itself; is either trivalent or quintivalent; and at a temperature of —13° C. and a pressure of 280 atmospheres becomes a colorless liquid.

The Atmosphere.

Atmospheric air is a mixture of nitrogen with oxygen and several other gases, of which carbon dioxide, water vapor, ammonia, nitric acid,

argon, and helium are important. Small quantities of sulphur dioxide, hydrogen sulphide, nitrous acid, and hydrocarbons, as well as suspended solid particles of organic (frequently bacteria) and inorganic materials, are frequently present.

Argon and **Helium,** elements discovered in 1894 and 1895, respectively, but previously included in analysis of air as nitrogen, resemble nitrogen in showing little affinity for other elements. The atomic weight of argon is 40, and that of helium 4.

THE AVERAGE COMPOSITION OF ATMOSPHERIC AIR BY VOLUME.

Name of Constituent.	*Volumetric Proportion.*
Nitrogen	77.16
Oxygen	20.60
Argon	0.80
Carbon dioxide	0.04
Aqueous vapor	from 0.5 to 1.40
Ammonia	trace
Nitric acid	trace
Helium	trace

Oxygen, nitrogen, carbon dioxide, and watery vapor are essential for animal and vegetable life. *The quantity of carbon dioxide is constantly increased by animal respiration, while that of oxygen is correspondingly diminished simultaneously; this excess of carbon dioxide is taken up by plants with a corresponding liberation of oxygen.* In this way the quantity of each constituent is balanced.

DEW-POINT.—The quantity of moisture in the air increases as the temperature rises and diminishes as it falls. The temperature at which air begins to deposit moisture, on cooling, is called the *dew-point.*

ANALYSIS OF AIR is made as follows:

Moisture and Carbon Dioxide.—Two U tubes, one containing dry calcium chloride and the other, pieces of potassium hydroxide, are weighed separately and joined together. Through the tubes a measured quantity of air is passed and the tubes reweighed. The calcium chloride retains the moisture and the potassium hydroxide the carbon dioxide, and the gain in weight of each tube shows their quantity.

Oxygen.—A graduated glass jar filled with air is inverted over a dish of mercury and a piece of phosphorus introduced, which combines with the oxygen present, and the consequent loss of volume is the volume of oxygen present.

Nitrogen.—The difference between the sum of the foregoing constituents and the total quantity of air is usually taken as the quantity of nitrogen present.

The oxygen in the atmosphere is free because all substances having affinity for it have combined with it, and left this quantity in excess. *Nitrogen occurs* free because of its slight affinity for other elements.

Liquid Air.

Preparation.—By bringing air in special apparatus to a temperature of about —191° C., produced by allowing air compressed under 2000 lbs. to the square inch to escape through a needle-point orifice. In thus expanding, the air absorbs heat and lowers the temperature in its vicinity. By this low temperature another portion of compressed air is cooled and allowed to escape in a similar manner, producing more intense cold. The temperature thus falls by stages to —191° C.

Physical Properties.—This pale-blue substance remains liquid in open vessels for a long time, due to the intense cold (—200° C.) of its evaporation. If evaporation is prevented, it speedily becomes gaseous, exerting great pressure on its container. Liquid nitrogen volatilizes more rapidly than oxygen, therefore a liquid containing 80 per cent. of oxygen may be obtained by allowing liquid air to evaporate. This liquid even at its low temperature is a powerful supporter of combustion.

Bodies immersed in liquid air show many physical alterations caused by the low temperature; many malleable and elastic substances, like lead and rubber, become exceedingly brittle. Changes of color and other properties also take place.

COMPOUNDS OF NITROGEN AND HYDROGEN

Ammonia.

M. F., NH_3; **M. W.**, 17; **G. F.**, $N\begin{cases} H \\ H. \\ H \end{cases}$

Occurrence.—Ammonia is formed in *nature* by decomposition of nitrogenous organic matter—*e. g.*, meat, blood, urine, etc.

Preparation.—1. By destructive distillation (see page 30); 2. By treating ammonium salts with sodium, potassium, or calcium hydroxide:

$$2NH_4Cl + Ca(OH)_2 = CaCl_2 + 2H_2O + 2NH_3.$$

3. In the manufacture of illuminating-gas by destructive distillation of coal, ammonia is formed from the nitrogen of the coal and absorbed in water, through which the gas is passed. This solution is the commercial source of ammonia.

Physical Properties.—This colorless gas, of pungent odor and disagreeable taste, burns in oxygen, forming water and nitrogen; is soluble in water, 700 volumes to 1 at standard temperature and pressure; liquefies without color at —40° C., under 7 atmospheres pressure, and crystallizes at —80° C.

Chemical Properties.—It is a strong alkali, neutralizing acids and forming salts, and with water forms ammonium hydroxide:

$$NH_3 + H_2O = NH_4OH.$$

PHARMACOPŒIAL PREPARATIONS.—*Aqua ammoniæ, water of ammonia, spirit of hartshorn* (U. S. P.), is a 10 per cent. by weight solution in water, of a specific gravity of 0.96, and the odor, taste, and reaction characteristic of the gas.

Aqua ammoniæ fortior, stronger water of ammonia (U. S. P.), specific gravity 0.90 contains 28 per cent.

Hydrazine-diamine.

M. F., N_2H_4; **M. W.**, 32; **G. F.**, $\begin{matrix} H \\ H \end{matrix}\!\!>\!N{-}N\!<\!\!\begin{matrix} H \\ H \end{matrix}$.

OCCURRENCE AND PROPERTIES.—This colorless gas, a liquid at ordinary temperatures and a crystalline solid at 0° C., is formed by complex organic decomposition, combines with oxygen from the air into nitrogen and water, with acids into salts, with water into an hydroxide, and has all of the characteristics of ammonia.

Hydroxylamine.

M. F., NH_2OH; **M. W.**, 33; **G. F.**, $\begin{matrix} H \\ H \end{matrix}\!\!>\!N{-}OH$.

PREPARATION.—By the action of nascent hydrogen on nitric acid:

$$HNO_3 + 6H = NH_2OH + 2H_2O.$$

PROPERTIES.—Known only in solution; with acids forms salts, which may be regarded as ammonium salts, with 1 atom of hydrogen replaced by hydroxyl:

$$NH_2OH + HCl = NH_3OHCl.$$

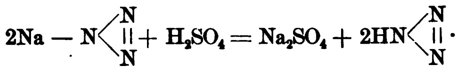
Triazoic Acid (Hydrazoic Acid).

M. F., N_3H; **M. W.**, 43; **G. F.**, $H{-}N\!<\!\!\begin{matrix} N \\ \| \\ N \end{matrix}$.

PROPERTIES.—This colorless gas, of disagreeable odor, poisons when inhaled, producing violent headache; is a strong acid, soluble in water, and forms insoluble compounds with silver and mercury.

PREPARATION.—1. From organic compounds; 2. By the action of *nitrogen* monoxide on the compound NH_2Na, formed by the action of *metallic sodium* on ammonia:

$$NH_3 + Na = NH_2Na + H.$$
$$NH_2Na + N_2O = NaN_3 + H_2O.$$



COMPOUNDS OF NITROGEN AND OXYGEN.

Oxides of Nitrogen.

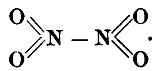
Nitrogen forms *five* oxides:

Nitrogen monoxide, $N_2O = N - O - N.$

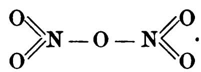
Nitrogen dioxide, $N_2O_2 = O = N - N = O.$

Nitrogen trioxide, $N_2O_3 = O = N - O - N = O.$

Nitrogen tetraoxide, $N_2O_4 =$



Nitrogen pentoxide, $N_2O_5 =$



STRUCTURE.—This is shown by the foregoing graphic formulas, in which the valence of nitrogen is either 1, 3, or 5.

Acids of Nitrogen.

PREPARATION.—The first, third, and fifth oxides with water form acids:

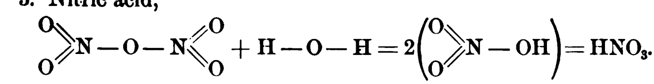
1. Hyponitrous acid,

$$N - O - N + H - O - H = 2(N - OH) = HNO.$$

2. Nitrous acid,

$$O = N - O - N = O + H - O - H = 2(O = N - OH) = HNO_2.$$

3. Nitric acid,



PROPERTIES.—Each acid, containing one atom of hydrogen replaceable by metals, is monobasic. Their salts are *hyponitrites*, *nitrites*, and *nitrates*. At a high temperature the molecules of nitrogen dioxide and

tetraoxide dissociate into NO and NO_2, unsaturated compounds. The terms nitrogen dioxide and tetraoxide are applied to them as well as to the saturated molecules.

Nitrogen Monoxide (Nitrous Oxide, Laughing Gas).

M. F., N_2O; **M. W.**, 44; **G. F.**, $N-O-N$.

PHYSICAL PROPERTIES.—It is a sweet, colorless, odorless gas, supports combustion, is soluble in equal volumes of cold water, and condenses into a colorless liquid under pressure of a 50 atmospheres. After inhalation in ascending quantities it produces exhilaration, and intoxication ("*laughing gas*"), then anæsthesia, (suitable for minor operations), and finally asphyxiation.

PREPARATION.—By heating *pure* ammonium nitrate:

$$NH_4NO_3 + \text{Heat} = 2H_2O + N_2O.$$

The ammonium nitrate must be pure. Chlorides are especially objectionable as impurities because gaseous compounds of chlorine and oxides of nitrogen may be formed.

The gas should be passed through two wash bottles, one of which contains NaOH, and the other $FeSO_4$ solution, in order to remove impurities.

Nitrogen Dioxide.

M. F., N_2O_2; **M. W.**, 60; **G. F.**, $O=N-N=O$.

PROPERTIES AND PREPARATION.—This colorless gas has a suffocating odor; is formed by the action of nitric acid on metals or reducing agents; combines with one or two atoms of oxygen into nitrogen tetraoxide:

$$2HNO_3 + 6H = 4H_2O + N_2O_2.$$

Nitrogen Tetraoxide.

M. F., N_2O_4; **M. W.**, 92 or 46; **G. F.**, $\begin{matrix}O\\O\end{matrix}\!\!>N-N<\!\!\begin{matrix}O\\O\end{matrix}$.

Formed as a *red* gas by absorption of oxygen when nitrogen dioxide is exposed to air. Red fumes are therefore observed whenever nitric acid acts on an oxidizing agent.

Nitrogen Trioxide.

M. F., N_2O_3; **M. W.**, 76; **G. F.**, $O=N-O-N=O$.

PROPERTIES.—This dark-blue liquid decomposes at —21° C. into NO *and NO_2, and with* water forms nitrous acid, HNO_2.

Nitrous Acid.

M. F., HNO_2; **M. W.,** 47; **G. F.,** $O = N - OH$.

The acid is not known free; the **nitrites,** its salts, are used extensively as reagents, and for other purposes.

Analytical Reactions of the Nitrites.

1. Give red fumes of NO_2 with sulphuric acid;
2. Decolorize solutions of $KMNO_4$;
3. Give a blue color with potassium iodide and starch-paste, due to liberated iodine acting on the starch;
4. Give a yellow to brown color with solution of meta-phenylene-diamine and sulphuric acid (see Water Analysis).

Nitrogen Pentoxide.

M. F., N_2O_5; **M. W.,** 108; **G. F.,** $\genfrac{}{}{0pt}{}{O}{O}\!\!>\!N-O-N\!<\!\!\genfrac{}{}{0pt}{}{O}{O}$.

Properties.—This white crystalline mass is easily decomposed by water into nitric acid.

Nitric Acid (Hydric Nitrate, Aqua Fortis).

M. F., HNO_3; **M. W.,** 63; **G. F.,** $\genfrac{}{}{0pt}{}{O}{O}\!\!>\!N-O-H$.

Varieties.—*Nitric Acid*, U. S. P., Sp. G. 1414, contains 68 per cent. nitric acid in water.

Dilute Nitric Acid, U. S. P., Sp. G. 1057, contains 10 per cent. nitric acid in water.

Fuming or *Commercial Nitric Acid*, Sp. G. from 1450 to 1560, is a brown liquid emitting fumes of NO_2, to which it owes its color.

Occurrence.—In nature, by the decomposition of nitrogenous organic matter, whose nitrogen is oxidized and unites with bases into nitrates: thus natural nitre-beds, as that in Chile, of sodium nitrate (*Chile salt-petre*) occur.

Preparation.—By the decomposition of Chile salt-petre by sulphuric acid:

$$NaNO_3 + H_2SO_4 = NaHSO_4 + HNO_3.$$

These proportions are used because of the greater solubility of sodium acid sulphate over sodium sulphate and the greater facility in cleansing the vessels.

PHYSICAL PROPERTIES.—This faintly yellow, fuming, corrosive liquid, with a suffocating odor, and a strong acid reaction, changes its color to yellow or yellowish red in sun-light by partial decomposition into nitrogen tetraoxide, water, and oxygen.

CHEMICAL PROPERTIES.—A strong oxidizing agent; volatilized by heat; turns vegetable and animal matter yellow, destroying the tissue.

Its *salts* are all soluble in water. In combining with metals, like copper and silver, red fumes of NO_2 are evolved, explicable in two ways:

1. The metal replaces the hydrogen of the acid directly, and the nascent hydrogen, a *strong reducing agent*, decomposes more nitric acid, a *strong oxidizing agent*, forming water and nitrogen dioxide—*i. e.*,

$$Ag + HNO_3 = AgNO_3 + H;$$

$$HNO_3 + 3H = 2H_2O + NO.$$

2. The nitric acid, by its strong oxidizing power, converts the metal into its oxide, and the oxide then unites with the acid to form the salt—*i. e.*,

$$3Cu + 2HNO_3 = 3CuO + H_2O + N_2O_2;$$

$$CuO + 2HNO_3 = Cu(NO_3)_2 + H_2O.$$

Oxygen-absorption by N_2O_2 and its effects have been explained.

TOXICOLOGY.—Nitric acid is a violent corrosive poison.

CHEMICAL ANTIDOTES.—Dilute alkalies, as solution of sodium carbonate, mixture of magnesia and water, and milk of lime.

Analytical Characters of Nitric Acid and the Nitrates.

1. Nitric acid, or nitrates and *sulphuric acid* heated with copper filings evolve *red* fumes of NO_2.

2. Nitric acid changes the color of *indigo* to *yellow;* nitrates and sulphuric acid, *heated*, produce this change, but not in the *cold*.

3. When a dilute nitrate-solution is mixed in a test-tube with a few drops of a solution of 1 part of *brucine* in 300 of 5 per cent. H_2SO_4, and then *sulphuric acid* is poured down the side of the tube so that the liquids form two layers, a *red* color, *fading to yellow*, develops at the contact-line. *This test shows one part of nitric acid in one million parts of water.*

4. When *diphenylamine* replaces the brucine in reaction 3, a *blue* color develops.

5. When *pyrogallic acid* replaces the brucine, a *brown* color is produced.

The diphenylamine test shows one part of nitric acid in three million, and pyrogallic acid shows one in ten million parts of water. The reagents themselves must be free from nitrates.

QUESTIONS.

Give the valence, atomic weight, and history of the discovery of nitrogen.

How may pure nitrogen be obtained?

How much nitrogen may be obtained from 100 lbs. of pure ammonium nitrite?

State the physical and chemical properties of nitrogen, argon, and helium.

State the composition of atmospheric air.

What constituents are necessary for animal and vegetable life, and how may their quantities be determined in air?

What is the dew-point?

What is the composition of ammonia, how is it formed, and what are its properties?

What are ammonium hydroxide, aqua ammonia, and stronger water of ammonia?

Give the composition and properties of hydrazine, hydroxylamine, and triazoic acid.

How many oxides of nitrogen are known, which form acids, and what is the valence of nitrogen in each?

What compounds are included in the terms nitrogen dioxide and tetraoxide?

What is meant by dissociation?

How is nitrogen monoxide made, and what are its properties?

How much pure ammonium nitrate will be required to make 132 lbs. of laughing-gas?

What change takes place when nitrogen dioxide is exposed to air?

Give tests for nitrites.

What is the composition of nitric acid, how is it made, and what are its properties?

How are nitrates formed in nature?

What is the strength of common nitric acid, and of dilute nitric acid?

What is fuming nitric acid?

What substance is constantly formed when nitric acid acts as an oxidizing agent?

Give tests for nitric acid and nitrates.

What chemical antidotes are to be used in nitric acid poisoning?

How much nitric acid may be obtained from 100 lbs. of Chile salt-petre?

CHAPTER VIII.

CARBON, ALLOTROPIC FORMS OF CARBON, AND COMPOUNDS OF CARBON.

CARBON.

S., C; **V.**, C_4^{iv}; **A. W.**, 12.

OCCURRENCE.—*Combined* in all organic matter, carbonates of metals, and carbon dioxide of the air, *free* and *crystallized* in diamond and graphite, *amorphous* and *impure* in lampblack, charcoal, etc.

ALLOTROPIC FORMS OF CARBON.

1. **Diamond** is pure carbon, in colorless crystals, the hardest of known substances, resists the action of reagents, when heated in the oxyhydrogen blowpipe burns, forming carbon dioxide.

2. **Graphite** crystallizes in six-sided prisms, of dark-gray color, soft, and greasy to the touch, leaves a black mark on white surfaces, resists heat and reagents like diamond, is used as a lubricator, for making stove-polish, lead-pencils, and, mixed with clay, crucibles, etc.

3. **Amorphous carbon** is a black solid whose hardness and other properties depend on its source. It is OBTAINED by incomplete combustion of organic matter, USED as *fuel*, *black pigment* (printers' ink), and as an *absorbent* of gases and other forms of matter, as in filters.

PHYSICAL PROPERTIES.—All forms of carbon are insoluble in ordinary solvents; in small quantity it may be dissolved in molten iron; on cooling most of it separates as graphite, and on cooling under pressure small crystals resembling diamonds are deposited. At a very high temperature (electric furnace) carbon becomes soft, and in small quantity is volatilized.

CHEMICAL PROPERTIES.—It has slight affinity for metals and oxygen at ordinary temperatures; at a very high temperature combines with many metals into carbides, and at red heat becomes a vigorous reducing agent, removing oxygen from oxidizing agents, or combining with free oxygen, usually indirectly.

COMPOUNDS OF CARBON AND OXYGEN.

Carbon Dioxide (Carbonic Anhydride).

M. F., CO_2; **M. W.**, 44; **G. F.**, $C\begin{matrix}\diagup\!\!\!\!=O\\ \diagdown\!\!\!\!=O\end{matrix}$.

PHYSICAL PROPERTIES.—This colorless, odorless gas, of acid taste is heavier than air (Sp. G. 1.524); liquefies without color under a pressure of 38 atmospheres at 0° C., at —79° C. forms white crystals; is soluble in its own volume of water at ordinary temperatures, and progressively soluble one volume for each atmosphere of increased pressure.

Soda water is its solution under about 10 atmospheres.

CHEMICAL PROPERTIES.—Carbon dioxide is not combustible and extinguishes flames; 10 per cent. by volume of it will fully check combustion. Although carbon dioxide is not poisonous through the stomach, it suffocates.

OCCURRENCE.—It occurs in nature in the air, spring-waters, and as carbonates, constantly forms during combustion of carbon compounds, organic decay, fermentation and putrefaction, and animal life-processes, being excreted by the lungs, skin, etc.

PREPARATION.—By the decomposition of carbonates by heat or acids:

$$CaCO_3 = CaO + CO_2;$$

$$CaCO_3 + H_2SO_4 = CaSO_4 + H_2O + CO_2.$$

Carbonic Acid.

M. F., H_2CO_3; **M. W.**, 62; **G. F.**, $O{=}C\begin{matrix}\diagup OH \\ \diagdown OH\end{matrix}$.

PREPARATION.—By the combination of carbon dioxide and water—*i. e.*,

$$CO_2 + H_2O = H_2CO_3$$

PROPERTIES.—This very weak dibasic acid forms salts (carbonates), all of which are insoluble except the alkali carbonates. The acid cannot be obtained pure, but only in dilute solutions, as in natural waters.

Tests for Carbonic Acid, Carbonates, and Carbon Dioxide.

1. Pass carbon dioxide through lime-water; a turbidity results by the formation of calcium carbonate, but, on continuing, passes away by the formation of soluble calcium bicarbonate. On boiling, carbon dioxide escapes and the turbidity returns.
2. Add to a carbonate any acid, collect the gas evolved and examine as above.
3. The addition of a solution of an alkaline carbonate precipitates solutions of other metals.

Carbon Monoxide (Carbonic Oxide).

M. F., CO; **M. W.**, 28; **G. F.**, $C{=}O$.

PROPERTIES.—This colorless, odorless, tasteless, neutral gas burns with the pale-blue flame frequently seen playing over a coal fire, forming CO_2, is very poisonous if inhaled, forming carbon-monoxide-hæmoglobin, which prevents the absorption of oxygen by the blood.

PREPARATION.—1. By passing carbon dioxide over red-hot coal; 2. By decomposition of oxalic acid by sulphuric acid; 3. By passing steam over red-hot coal (carbon), as in the manufacture of *water gas*, which is used for heating, and after it has been mixed with hydrocarbons for illumination.

COMPOUNDS OF CARBON, HYDROGEN, AND NITROGEN.

These will be considered under Organic Chemistry.

Combustion and Its Products.

CHEMICAL BASIS.—*Flame* is vapor undergoing combustion, which usually forms gases: *hydrogen*, *carbon monoxide*, methane, and ethane.

Methane and *ethane* are the compounds of carbon and hydrogen, CH_4 and C_2H_4, generated during *organic combustion* or by *destructive distillation* and then burned.

In a burning candle the fat is decomposed by the heat of the flame into the gases, which burn until all of the fat and gases become carbon dioxide and water.

The Ordinary Flame.—CHEMICAL PARTS.—*Three:* 1. The dark innermost portion of volatilized combustible matter; 2. A surrounding luminous cone; 3. An outermost almost non-luminous cone, having the highest temperature of the flame. A cold surface held in the middle cone will be covered by a carbon-deposit (soot).

FIG. 18.

Structure of flame. (Simon.)

EXPLANATION OF THE THREE CONES.—The affinity of hydrogen for oxygen is greater than that of carbon, and in the outer colorless envelope sufficient oxygen is supplied by the air for complete combustion of the gases. Only a limited quantity of oxygen penetrates the outer envelope, and this is taken up by the hydrogen of the hydrocarbons. The liberated carbon is heated intensely in the burning hydrogen and gives the luminosity to the flame. On reaching the outer cone the carbon finds sufficient oxygen for complete combustion. If sufficient air be mixed with the gas before ignition or be forcibly driven into the flame, no deposit of carbon occurs, and the illuminating power almost disappears, but a more intense heat is generated.

FIG. 19.

The **Bunsen burner**, the **blow-pipe**, and the **blast furnace** are constructed on the above principle to furnish intense heat to the flame.

FLAME OF THE BLOW-PIPE (Fig. 19) is peculiar, having two long pointed cones, the outer yellowish and the inner blue. The space between these cones is filled with highly heated combustible material of strong reducing properties, and called the *reducing flame.* The highly heated air just beyond the tip of the outer cone is called the *oxidizing flame.*

BLAST-FURNACE FLAME.—The blast furnace is mechanically so constructed that air is rapidly introduced into the burning mass, thus increasing the rate of consumption of the fuel and raising the heat to an intense degree. The structure of the blast-furnace flame resembles that of the blow-pipe.

OXYHYDROGEN FLAME is formed by burning hydrogen in pure oxygen, and has an extremely high temperature, and is solid, combustion *occurring simultaneously* throughout it.

QUESTIONS.

How is carbon found in nature, and what are the properties of its allotropic modifications?

State composition, occurrence in nature, mode of manufacture, and properties of carbon dioxide.

What is carbon monoxide, how is it formed, and what are its properties?

State composition, properties, and tests for carbonic acid.

What is meant by flame?

What gases are generated during the combustion of ordinary organic matter?

Describe the structure of a flame, and give reasons for this structure.

What is the principle of the construction of a Bunsen burner?

What is the oxyhydrogen flame?

CHAPTER IX.

BORON AND SILICON, AND THEIR COMPOUNDS.

BORON.

S., B; **V.**, 3; **A. W.**, 10.9.

OCCURRENCE.—*In nature* in small quantities as boric acid or borax (*sodium borate*) in certain lakes in Italy, in Nevada, and in California.

PREPARATION.—1. As amorphous greenish-brown *powder*, by decomposing borax trioxide with sodium.

2. As *crystals*, by heating the amorphous powder with aluminum.

CHEMICAL PROPERTIES.—Boron combines with many non-metals —*e. g.*, chlorine as BCl_3, fluorine as BFl_3, and hydrogen as BH_3; directly with nitrogen at a high temperature, as nitrogen boride BN.

COMPOUNDS OF BORON.

Borax (Sodium Borate, Sodium Biborate).

M. F., $Na_2B_4O_7.10H_2O$; **M. W.**, 381.6; **G. F.**, $B_4O_5\langle{}^{ONa}_{ONa}$.

OCCURRENCE, PREPARATION.—This sodium salt of triboric acid is formed in nature, and by recrystallizing is refined, containing then ten molecules of water of crystallization.

PHYSICAL PROPERTIES.—It is colorless, slightly efflorescent, soluble in water and in an equal weight of glycerin at 80° C., insoluble in alcohol.

CHEMICAL PROPERTIES.—By heat borax is fused to a colorless liquid which cools to a transparent mass (*borax glass*). Fused borax combines

with certain metallic oxides as double borates, with characteristic colors. It is therefore used in blow-pipe analysis and the manufacture of colored glazes and enamels.

Boric Acid (Boracic Acid).

M. F., H_3BO_3; **M. W.**, 61.9; **G. F.**, $\mathrm{B}\begin{array}{l}\diagup \mathrm{OH} \\ - \mathrm{OH}. \\ \diagdown \mathrm{OH}\end{array}$

Occurrence.—In small quantities in nature in the pure state.

Preparation.—By adding hydrochloric acid to hot concentrated solution of borax; boric acid separates on cooling:

$$Na_2B_4O_7 + 2HCl + 5H_2O = 4H_3BO_3 + 2NaCl.$$

Properties.—This weak acid is a white crystalline substance, sparingly soluble in alcohol and water.

Chemical Properties.—By heat it becomes successively metaboric acid (HBO_2), tetraboric acid ($H_2B_4O_7$), and finally at white heat boron trioxide (B_2O_3). Boric acid and borax are used as antiseptics and preservatives.

Qualitative Analysis for Boric Acid and the Borates.

1. A concentrated neutral solution of a borate forms a white precipitate when added to a neutral solution of a calcium, barium, or silver salt. The precipitate formed will be calcium, barium, or silver borate.
2. Borax or boric acid solution turns turmeric paper brown.
3. A mixture of a borate, sulphuric acid, and alcohol burns with a flame having a mantle of green color.
4. Borax fused on a loop of platinum wire forms a colorless bead (*borax bead*) which when moistened with sulphuric acid colors the flame of the Bunsen burner green.

SILICON.

S., Si; **V.**, 4; **A. W.**, 28.3.

Occurrence.—In large quantity as *silicon dioxide* (*silica*), SO_2, in sand, rocks, crystals, quartz, and sandstones, and as *silicates* in rocks, as granite, slate, mica, etc.; in small quantities in spring-waters, derived from the soil, and in animal and vegetable matter.

Properties.—Silicon resembles carbon, is quadrivalent, amorphous, and dimorphous.

COMPOUNDS OF SILICON.

Silicon Hydride.—

M. F., SiH_4; **M. W.**, 32.3; **G. F.**, $\mathrm{Si}\begin{array}{l}\diagup \mathrm{H} \\ < \mathrm{H} \\ \phantom{<} \mathrm{H}. \\ \diagdown \mathrm{H}\end{array}$

Silicon Chloride.—

$$\textbf{M. F.},\ SiCl_4;\ \textbf{M. W.},\ 169.9;\ \textbf{G. F.},\ Si{<}\begin{matrix}Cl\\Cl\\Cl\\Cl\end{matrix}.$$

Silicon Fluoride.—

$$\textbf{M. F.},\ SiFl_4;\ \textbf{M. W.},\ 104.3;\ \textbf{G. F.},\ Si{<}\begin{matrix}Fl\\Fl\\Fl\\Fl\end{matrix}.$$

All are gases, formed by direct combination, except $SiCl_4$, which is a *liquid*, and SiH_4, which is formed by electrolysis of NaCl, with aluminum and silicon electrodes.

Silicon forms two acids—*i. e.*,

Normal silicic acid,

$$\textbf{M. F.},\ H_4SiO_4;\ \textbf{M. W.},\ 96.3;\ \textbf{G. F.},\ Si{<}\begin{matrix}OH\\OH\\OH\\OH\end{matrix}.$$

Ordinary silicic acid,

$$\textbf{M. F.},\ H_2SiO_3;\ \textbf{M. W.},\ 78.3;\ \textbf{G. F.},\ SiO{<}\begin{matrix}OH\\OH\end{matrix}.$$

The latter decomposes by heat into water and silicon dioxide.

Silicates usually are complex salts of acids formed by the condensation of two or more molecules of normal silicic acid, with removal of one or more molecules of water. They are insoluble in ordinary solvents, in analysis they are decomposed and rendered soluble by the action of hydrofluoric acid, and converted into silicon fluoride, $SiFl_4$. Water decomposes this compound, forming silicic acid and hydrofluosilicic acid—*i. e.*:

$$3SiF_4 + 3H_2O = H_2SiO_3 + 2H_2SiF_6.$$

Silicon Carbide (Carborundum, Carbon Silicide).—

$$\textbf{M. F.},\ SiC;\ \textbf{M. W.},\ 40.3;\ \textbf{G. F.},\ Si{\equiv}C.$$

PROPERTIES.—A beautiful, dark-green crystalline material, of hardness second to the diamond, used as a polishing agent.

PREPARATION.—By heating sand (SiO_2) and carbon with sodium chloride in an electric furnace to 3500° C., when the carbon combines with both constituents of the sand:

$$SiO_2 + 3C = 2CO + SiC.$$

Glass, is a *mixture* of several silicates, chiefly of sodium, calcium, and lead, with an excess of silica.

COLORED GLASS is formed by fusing certain metallic oxides with these silicates—*e. g.*, *cobalt* produces a *blue* color; *manganese*, an *amethyst*; *ferric oxide*, a *brownish-yellow glass*.

SOLUBLE GLASS is a *mixture* of alkaline silicates, with an excess of alkaline hydroxide, obtained by fusing silica with an excess of alkaline carbonates or hydroxide.

QUESTIONS.

State symbols and valence of boron and silicon.
How does boron occur in nature, and how may the element be obtained from its compounds?
What peculiar property is shown by boron with regard to nitrogen?
State composition and properties of borax.
How is boric acid manufactured, and what are its properties?
Give tests for borates.
How does silicon occur in nature, and how does the element resemble carbon?
What acids are formed by silicon?
How does hydrofluoric acid act on silicates?
State composition, mode of manufacture, and properties of carborundum.
What is glass? Colored glass? Soluble glass?

CHAPTER X.

SULPHUR (SELENIUM AND TELLURIUM).

S., S; **V.**, 2, 4, or 6; **A. W.**, 32.

OCCURRENCE.—In nature, *free*, in volcanic districts, as Sicily, *combined* as sulphates, chiefly gypsum ($CaSO_4.2H_2O$), also as sulphides, in pyrites (FeS_2), cinnabar (HgS), galena (PbS), etc.: in small quantity in animal and vegetable matter, and as hydrogen sulphide in spring-waters.

PHYSICAL PROPERTIES.—This yellow, brittle, odorless, tasteless solid is soluble in benzene, ether, chloroform, carbon disulphide, essential and fat-oils, insoluble in water and alcohol, melts at 115° C. to a thin liquid, on further heating gradually thickens, until at 200° C. it scarcely flows, above this point gradually thins again, and boils at about 440° C.

CHEMICAL PROPERTIES.—Sulphur resembles oxygen, supports combustion in the gaseous state, and forms many analogous compounds—*e. g.*, H_2S, H_2O; CS_2 and CO_2.

VARIETIES OF SULPHUR AND THEIR PREPARATION.

Crude sulphur. By heating native sulphur, and separating the melted sulphur from the impurities.

Sublimed sulphur (*flowers of sulphur*), U. S. P. By subliming crude sulphur in suitable apparatus.

Roll sulphur is sublimed sulphur melted and poured into moulds. *Rubbed with* hair or wool it generates negative electricity.

Washed sulphur (*sulphur lotum*), U. S. P. By washing sublimed sulphur with dilute ammonia, and then water to remove acid impurities formed during the sublimation and to free it from arsenic.

Precipitated sulphur (*milk of sulphur*), U. S. P. By boiling 1 part of calcium hydroxide, 2 parts of sulphur, and 30 parts of water, filtering, adding hydrochloric acid to the filtrate, and washing and drying the precipitate.

ALLOTROPIC FORMS OF SULPHUR.

Amorphous Sulphur.

This is a plastic mass obtained by pouring boiling sulphur into cold water.

Polymorphous Sulphur.

From solution in carbon disulphide, sulphur crystallizes into octahedra, its natural crystal; when melted and cooled, into six-sided prisms.

Qualitative Analysis for Sulphur.

HEPAR TEST.—*This is applicable to sulphur in any combination.* A mixture of the suspected material with potassium cyanide and sodium carbonate is fused on charcoal in a blow-pipe flame, placed on a clean surface of silver, and moistened with hydrochloric acid. A black stain forms when sulphur is present.

COMPOUNDS OF SULPHUR AND HYDROGEN.

Hydrogen Sulphide (Hydrosulphuric Acid, Sulphuretted Hydrogen).

M. F., H_2S; **M. W.,** 34; **G. F.,** $\begin{matrix} H \\ H \end{matrix} \!\! > \! S$.

PHYSICAL PROPERTIES.—This colorless gas, of disagreeable odor and taste, and directly poisonous when inhaled, is soluble in water (three volumes to one), forming a feebly acid solution, which is used as a reagent.

CHEMICAL PROPERTIES.—The gas burns with a blue flame, forming sulphur dioxide and water. Forms insoluble compounds with heavy metals.

OCCURRENCE.—In nature, by putrefaction of animal and vegetable matter, and in certain mineral waters.

PREPARATION.—By the decomposition of ferrous sulphide by dilute sulphuric acid—*i. e.*,

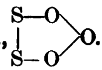
$$FeS + H_2SO_4 = FeSO_4 + H_2S.$$

USES.—As a reagent in analytical chemistry. Its value depends on the insoluble compounds of suphur with many metals, many having characteristic color.

COMPOUNDS OF SULPHUR AND OXYGEN.

Oxides of Sulphur.

1. Sulphur dioxide (sulphurous anhydride).
2. Sulphur sesquioxide:

M. F.,[1] S_2O_3; **M. W.**, 112; **G. F.**, .

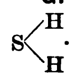
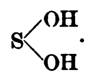
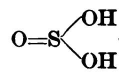
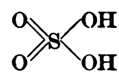
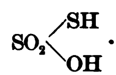
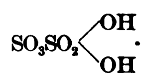
3. Sulphur trioxide (sulphuric anhydride).
4. Persulphuric anhydride;

M. F., S_2O_7; **M. W.**, 176; **G. F.**, not determined.

The first and third only are important.

Acids of Sulphur.

Eleven are known (some only as constituents of salts):

Common Name.	M. F.	M. W.	G. F.
1. Hydrosulphuric acid	H_2S	34	
2. Hyposulphurous acid	H_2SO_2	66	
3. Sulphurous acid	H_2SO_3	82	
4. Sulphuric acid	H_2SO_4	98	
5. Thiosulphuric acid	$H_2S_2O_3$	114	
6. Pyrosulphuric acid	$H_2S_2O_7$	178	

[1] The M. F., M. W., and G. F. of compounds are given in these lists only **when** *they are not subsequently* described in the text.

Common Name.	M. F.	M. W.	G. F.
7. Persulphuric acid	$H_2S_2O_8$	194	
8. Dithionic acid	$H_2S_2O_6$	162	
9. Trithionic acid	$H_2S_3O_6$	194	Not determined.
10. Tetrathionic acid	$H_2S_4O_6$	226	
11 Pentathionic acid	$H_2S_5O_6$	258	

Sulphur Dioxide (Sulphurous Anhydride).

M. F., SO_2; **M. W.**, 64; **G. F.**, $S{\langle}{O \atop O}$.

PHYSICAL PROPERTIES.—This colorless gas is of suffocating odor, poisonous when inhaled pure, and irritating when diluted with air; at —10° C. liquefies, and at —75° C. solidifies. About forty volumes of the gas dissolve in one of water.

CHEMICAL PROPERTIES.—It is a strong deoxidizing and bleaching agent, and *a good disinfectant for dwellings, etc., provided it is generated by burning four pounds of crude sulphur to each thousand cubic feet of air-space* (Park).

PREPARATION.—1. By burning sulphur or certain sulphides in air or pure oxygen. 2. By the action of sulphuric acid on certain metals of the silver and lead group, or on charcoal—*i. e.*,

$$2H_2SO_4 + Cu = CuSO_4 + 2H_2O + SO_2;$$
$$2H_2SO_4 + C = CO_2 + 2H_2O + 2SO_2.$$

Sulphurous Acid.

PREPARATION.—1. By combining sulphur dioxide and water:

$$SO_2 + H_2O = H_2SO_3:$$

2. *Sulphurous acid*, U. S. P., containing 6.4 per cent. by weight of sulphur dioxide, is made by the action of charcoal on sulphuric acid, the generated gas being dissolved in water.

PHYSICAL AND CHEMICAL PROPERTIES.—This colorless liquid, completely volatilized by heat, has the odor of sulphur dioxide, and deoxi-

dizing, bleaching, and disinfecting properties, is dibasic, and forms salts, called *sulphites*. The acid and the sulphites absorb oxygen and form sulphuric acid and sulphates.

Qualitative Analysis for Sulphurous Acid and Sulphites.

Add their solution to one of potassium permanganate acidified with hydrochloric acid and boil. The potassium permanganate will be decolorized. Then test for sulphuric acid in the solution.

Sulphur Trioxide (Sulphuric Anhydride).

$$\textbf{M. F.}, SO_3;\ \textbf{M. W.}, 80;\ \textbf{G. F.}, O{=}S\begin{matrix}\diagup O \\ \diagdown O.\end{matrix}$$

PROPERTIES AND PREPARATION.—This white silky crystalline substance, obtained by simultaneously passing oxygen and sulphur dioxide over red-hot spongy platinum, is used chiefly in manufacturing fuming sulphuric acid.

Sulphuric Acid (Oil of Vitriol).

PREPARATION.—By simultaneously passing into leaden chambers vapors of sulphur dioxide from burning sulphur or pyrites, nitric acid, steam, and air. The sulphur dioxide combines with water and nitric acid into sulphuric acid and nitrogen trioxide. The nitrogen trioxide combines with sulphur dioxide, oxygen, and water, forming nitrosyl-sulphuric acid, which is split by the steam into sulphuric acid and nitrogen trioxide. This nitrogen trioxide forms more nitrosyl-sulphuric acid, which is immediately decomposed as above. *Apparently a given quantity of nitric acid should form an unlimited quantity of sulphuric acid, but the loss in the oxides of nitrogen during the process causes gradual disappearance of the nitric acid.* These reactions are:

1. $2SO_2 + H_2O + 2HNO_3 = 2H_2SO_4 + N_2O_3$;

2. $N_2O_3 + 2SO_2 + 2O + H_2O = 2(NO_2SO_2OH)$;

3. $2(NO_2SO_2OH) + H_2O = 2H_2SO_4 + N_2O_3$.

VARIETIES.—The sulphuric acid, with some water, collects at the chamber-bottom. *Chamber acid*, Sp. G. 1.55, is drawn off and concentrated by evaporation, first in leaden pans and afterward in platinum dishes.

Pan sulphuric acid (Sp. G. 1.75, 78 per cent. pure H_2SO_4) is the *strongest* obtainable by evaporation in lead pans, as more concentrated acid dissolves lead.

Concentrated sulphuric acid (Sp. G. 1.84) contains 95 per cent. pure H_2SO_4.

Dilute sulphuric acid, U. S. P., contains 10 per cent. H_2SO_4.

Sulphuric acid was originally manufactured by distillation from ferrous sulphate (green vitriol); because of the oily appearance of the distillate, it was called *oil of vitriol*, its present common name.

PHYSICAL PROPERTIES.—Concentrated sulphuric acid is a colorless oily liquid, boils at 338° C. and freezes at —26° C.

CHEMICAL PROPERTIES.—This strong dibasic acid forms salts called *sulphates*, by combinations with metals or bases, or by displacing other acids from their salts. It absorbs water, and is therefore a desiccating agent, destroys and chars organic matter by removing hydrogen and oxygen in the proportion to form water, and depositing carbon. Its slough in animal tissue is, consequently, black.

Analytical Characters of Sulphuric Acid or Sulphates.

1. Addition of barium chloride precipitates barium sulphate, white and insoluble in acids.
2. Addition of solution of lead acetate precipitates lead sulphate.

Pyrosulphuric Acid (Disulphuric Acid, Fuming or Nordhausen Sulphuric Acid).

PROPERTIES.—This brown oily, very corrosive liquid fumes in the air, and decomposes into SO_3 and H_2SO_4 when heated.

PREPARATION.—By dissolving SO_3 (obtained by distilling $FeSO_4$) in sulphuric acid—*i. e.*,

$$H_2SO_4 + SO_3 = H_2S_2O_7$$

Thiosulphuric Acid.

PROPERTIES.—Unknown in the free state.

Sodium Thiosulphate (Sodium Hyposulphite, U. S. P.).

M. F., $Na_2S_2O_3$; **M. W.**, 138; **G. F.**, $SO_2\begin{cases} SNa \\ ONa \end{cases}$

PROPERTIES.—This salt of thiosulphuric acid decomposes by acids, liberating sulphur and sulphur dioxide, with silver nitrate forms silver thiosulphate, soluble in an excess of the sodium salt, forming a double compound, $AgNaS_2O_3$. This property is used in photography for removing unchanged silver from the plate after developing.

OTHER COMPOUNDS OF SULPHUR.

Carbon Disulphide (Carbon Bisulphide).

M. F., CS_2; **M. W.**, 76; **G. F.**, $\mathrm{C}\begin{smallmatrix} =\mathrm{S} \\ =\mathrm{S} \end{smallmatrix}$

PROPERTIES.—This colorless, volatile, inflammable, nearly neutral liquid has high refractive power, Sp. G. 1.29, boils at 460° C., and is used as a solvent.

PREPARATION.—By passing vapor of sulphur dioxide over heated charcoal.

SELENIUM AND TELLURIUM.

S., Se, Te; **A. W.**, 79, 125; **V.**, 2, 4 or 6, respectively.

PROPERTIES.—These elements resemble sulphur, are polymorphous, and combine with hydrogen and oxygen similarly to sulphur—*e. g.*,

H_2S	H_2Se	H_2Te
SO_2	SeO_2	TeO_2
SO_3	SeO_3	TeO_3
H_2SO_3	H_2SeO_3	H_2TeO_3
H_2SO_4	H_2SeO_4	H_2TeO_4

QUESTIONS.

How is sulphur found in nature?

What is meant by crude sulphur, flowers of sulphur, washed sulphur, and precipitated sulphur, and how are they obtained?

State physical and chemical properties of sulphur.

State composition, mode of manufacture, properties, and uses of hydrogen sulphide, sulphur dioxide, sulphurous acid, and sulphuric acid.

How may sulphites be distinguished from sulphates?

Give tests for sulphates.

State strength of concentrated and dilute sulphuric acid.

What is the composition of pyrosulphuric acid?

State composition and uses of sodium thiosulphate.

How may the presence of sulphur in a compound be demonstrated?

How do selenium and tellurium resemble sulphur?

How much U. S. P. sulphurous acid can be obtained from 100 grams of pure sulphuric acid?

CHAPTER XI.

PHOSPHORUS.

S., P; **V.**, 3 or 5; **A. W.**, 31.

HISTORY.—Discovered and made from urine in 1669 by Brandt, of Hamburg.

OCCURRENCE.—In nature, as phosphates of calcium, iron, aluminum, and other metals in rocks and the soil. From the soil it is taken up by plants, and thence enters the animal body in organic and inorganic combination. Bone contains about 50 per cent. of calcium phosphates. Phosphorus is eliminated in the urine.

PREPARATION.—By heating a dry mixture of metaphosphoric acid and charcoal in a retort and collecting the products under water—*i. e.*,

$$2HPO_3 + 5C = H_2O + 5CO + 2P.$$

The metaphosphoric acid is obtained from natural calcium phosphate by the action of sulphuric acid, and the heating of the orthophosphoric acid so obtained at a strong red heat—*i. e.*,

$$Ca_3(PO_4)_2 + 3H_2SO_4 = 3CaSO_4 + 2H_3PO_4;$$
$$H_3PO_4 = H_2O + HPO_3.$$

During the reaction several intermediate compounds are formed.

PHYSICAL PROPERTIES.—This colorless, translucent, waxy solid becomes opaque, yellow, and yellowish red on standing, especially in sunlight, and brittle at 0° C., melts at 44° C., and volatilizes at 290° C., is insoluble in water, slightly soluble in alcohol, ether, fixed and essential oils, very soluble in chloroform and in bisulphide of carbon. The last solution is dangerously combustible, and from it phosphorus crystallizes.

CHEMICAL PROPERTIES.—Phosphorus has luminosity and great affinity for oxygen; when exposed to the air combines with oxygen, giving off a garlic-like odor and white vapors which are luminous in the dark. The heat of this slow oxidation soon reaches 50° C., when the phosphorus burns with a bright flame, forming white fumes of phosphoric oxide. Phosphorus is therefore kept under water, and manipulated away from oxygen. Phosphorus combines directly with many metals, forming *phosphides*, as zinc phosphide (Zn_3P_2), also with chlorine, bromine, iodine, and forms six compounds with sulphur. Phosphorus differs from other elements in containing four atoms in each molecule (MOLECULAR WEIGHT 124).

Phosphorus shows many points of similarity to nitrogen, in being three or five valent, and in uniting with hydrogen, oxygen, and chlorine into analogous compounds—*i. e.*,

NH_3	N_2O_3	N_2O_5	NCl_3
PH_3	P_2O_3	P_2O_5	PCl_3

Toxicology of Phosphorus.

CHRONIC POISONING occurs in workers in phosphorus.

ACUTE POISONING follows the ingestion of a poisonous quantity.

CHEMICAL ANTIDOTES.—1. Potassium permanganate, which converts phosphorus into phosphoric acid, by oxidation. 2. Oil of turpentine, whose action has not been explained. These should be preceded by emesis and lavage. *Milk and fats should not be given, because they dissolve phosphorus and assist its assimilation.*

TESTS FOR PHOSPHORUS IN THE STOMACH-CONTENTS:

1. By simple agitation in a dark room for the characteristic luminosity.

2. When the quantity of phosphorus is small this simple test fails. Then distil in a dark room a mixture of the stomach-contents, dilute sulphuric acid, and water. Phosphorus will distil over with the steam, and be seen as a luminous ring in the upper part of the condenser.

3. A minute quantity may have become oxidized. Therefore phosphorous acid must be tested for in the distillate.

ALLOTROPIC FORMS OF PHOSPHORUS.

The red or amorphous, and the black or metallic forms are of interest among several.

Red Phosphorus.

to 260° C. for about thirty-six hours, differs from ordinary phosphorus in being non-poisonous, non-luminous insoluble in solvents just mentioned, and non-combustible below 280° C., at which it becomes the ordinary form.

PROPERTIES AND PREPARATION.—This brick-red, amorphous powder, obtained by heating common phosphorus in an atmosphere of CO_2

Black or Metallic Phosphorus.

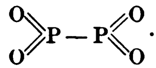
PREPARATION AND PROPERTIES.—Obtained by heating red phosphorus in a sealed tube to 500° C.; it is deposited in the cooler portion of the tube in microscopic needle-shaped crystals having metallic lustre.

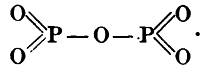
USES OF ELEMENTARY PHOSPHORUS.—In the manufacture of phosphorated oil, pills, and spirit of phosphorus, phosphoric acid, and other compounds, and chiefly (both common and red phosphorus) of matches.

COMPOUNDS OF PHOSPHORUS AND OXYGEN.

Oxides of Phosphorus.

Common Name.	M. F.	M. W.	G. F.
1. Phosphorus monoxide	P_4O	140	Not determined.
2. *Phosphorus trioxide*	P_2O_3	110	O=P—O—P=O.

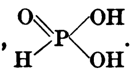
3. Phosphorus tetraoxide . . . . P_2O_4 124 

4. Phosphorus pentoxide . . . . P_2O_5 142 

PROPERTIES.—The pentoxide is formed by *rapid*, the others by *slow* oxidation. The trioxide and pentoxide unite with water into phosphorous and phosphoric acids.

ACIDS OF PHOSPHORUS.

Phosphorous Acid.

S., H_3PO_3; **M. W.**, 82; **G. F.**, 

PREPARATION.—By dissolving phosphorous oxide in water—*i. e.*,

$$P_2O_3 + H_2O = 2H_3PO_3.$$

PHYSICAL AND CHEMICAL PROPERTIES.—This colorless, liquid, strongly dibasic acid and its salts (*phosphites*) are strong reducing agents, uniting with oxygen into phosphoric acid and phosphates.

Analytical Characters of Phosphorous Acid and the Phosphites.

They reduce mercuric chloride, silver nitrate, and nitric acid, and phosphoric acid may be detected in the solution after this reaction is complete.

Phosphoric Acids.

Phosphoric oxide unites with water in three proportions—*i. e.*:

1. Metaphosphoric acid:

$$P_2O_5 + H_2O = 2HPO_3;$$

2. Pyrophosphoric acid:

$$P_2O_5 + 2H_2O = H_4P_2O_7;$$

3. Orthophosphoric acid:

$$P_2O_5 + 3H_2O = 2H_3PO_4.$$

Orthophosphoric Acid.

M. F., H_3PO_4; **M. W.**, 98; **G. F.**, $O=P\begin{cases}OH\\OH.\\OH\end{cases}$

PROPERTIES.—This colorless, odorless, strongly acid liquid is tribasic, forming three series of salts according to the number of hydrogen atoms replaced—*i. e.*,

1. Sodium-dihydrogen phosphate or primary sodium phosphate;

M. F., NaH_2PO_4; **M. W.**, 120; **G. F.**, $O=P\begin{cases}ONa\\OH.\\OH\end{cases}$

2. Disodium-hydrogen phosphate or secondary sodium phosphate;

M. F., Na_2HPO_4; **M. W.**, 142; **G. F.**, $O=P\begin{cases}ONa\\ONa.\\OH\end{cases}$

3. Trisodium phosphate or tertiary sodium phosphate;

M. F., Na_3PO_4; **M. W.**, 164; **G. F.**, $O=P\begin{cases}ONa\\ONa.\\ONa\end{cases}$

All alkali phosphates, but only the primary phosphates of other metals, are soluble in water.

PREPARATION.—1. Natural phosphates are salts of orthophosphoric acid, whence it may be obtained by sulphuric acid (see page 119).

2. By dissolving phosphoric oxide in water and boiling.

3. By treating phosphorus with dilute nitric acid, and evaporating the solution until any excess of nitric acid has been expelled.

VARIETIES.—*Orthophosphoric acid*, (*phosphoric acid*, U. S. P.), Sp. G. 1.710, contains 85 per cent. of pure acid.

Dilute phosphoric acid, U. S. P., contains 10 per cent. of pure acid.

Qualitative Analysis of Phosphoric Acid and the Phosphates.

1. A neutral solution of a phosphate precipitates solutions of barium, calcium, or ferric chloride, forming the respective phosphates, which are soluble in acid.

2. Silver nitrate forms with phosphates a yellow precipitate soluble in ammonia and in nitric acid.

3. A solution of ammonium molybdate in dilute nitric acid, when boiled with phosphates in an acid solution, forms a yellow precipitate of ammonium phosphomolybdate ($(NH_4)_3PO_4.10MoO_3.2H_2O$), soluble in ammonia water.

4. A mixture of magnesium sulphate, ammonium chloride, and *ammonia water* added to a neutral or alkaline solution of a phosphate

forms a white crystalline mass of ammonium-magnesium phosphate ($NH_4.MgPO_4$) soluble in acids.

Quantitative Analysis of Phosphoric Acid and the Phosphates.

The weighed quantity of material is dissolved in nitric acid, and an excess of ammonium molybdate added to the clear diluted solution. The yellow precipitate is dissolved in dilute ammonia and precipitated by magnesium-ammonia mixture (4). The washed precipitate is burned, forming magnesium pyrophosphate ($Mg_2P_2O_7$), from whose weight the quantity of phosphoric acid is calculated.

Pyrophosphoric Acid.

M. F., $H_4P_2O_7$; **M. W.**, 178; **G. F.**, $\begin{matrix} HO \\ O{=} \\ HO \end{matrix}\!\!>\!P{-}O{-}P\!<\!\!\begin{matrix} OH \\ {=}O \\ OH \end{matrix}$.

Properties and Preparation.—This tetrabasic acid, whose salts are called *pyrophosphates*, is formed by heating orthophosphoric acid to 200° C., is not precipitated by ammonium molybdate, and does not coagulate albumin.

Metaphosphoric Acid (Glacial Phosphoric Acid).

M. F., HPO_3; **M. W.**, 80; **G. F.**, $\begin{matrix} O \\ O \end{matrix}\!\!\gg\!P{-}OH.$

Properties and Preparation.—This glassy white solid, soluble in water, is formed: 1. By heating pyrophosphoric acid or ammonium phosphate; 2. By dissolving phosphoric oxide in water; and 3. Becomes orthophosphoric acid when the solution is boiled, and is again formed when orthophosphoric acid is heated to red heat. It is a monobasic acid, and differs from common phosphoric acid, by forming a white precipitate with ammonio-silver nitrate, coagulating albumin, and being poisonous.

Hypophosphorous Acid.

M. F., H_3PO_2; **M. W.**, 66; **G. F.**, $\begin{matrix} O \\ H \end{matrix}\!\!>\!P\!<\!\!\begin{matrix} H \\ OH \end{matrix}$.

Properties and Preparation.—This colorless, syrupy, strongly acid liquid, obtained by decomposing barium hypophosphite with dilute sulphuric acid, and concentrating the solution by evaporation in a vacuum, is unstable, decomposing by heat or by standing into phosphorus and

phosphoric acid. Hypophosphorous acid and its salts, the hypophosphites, are strong reducing agents, which must be remembered when prescribing them with mercuric and ferric salts.

STRUCTURE.—Hypophosphorous and phosphorous acids are peculiar; each contains three hydrogen atoms in the molecule, of which one in the former and two in the latter acid are replaceable by bases. It is assumed that the phosphorus in each is *quantivalent*, and that the replaceable hydrogen is joined to the phosphorus by an oxygen atom as *hydroxyl*.

VARIETIES.—*Dilute hypophosphorous acid*, U. S. P., contains 10 per cent. of pure acid.

The Hypophosphites.

PREPARATION.—By boiling phosphorus with hydroxides of the alkalies or alkaline earths:

$$3Ba(OH)_2 + 8P + 6H_2O = 3Ba(PH_2O_2)_2 + PH_3.$$

The phosphine (PH_3) formed ignites spontaneously.

COMPOUNDS OF PHOSPHORUS AND HYDROGEN.

Phosphine.

M. F., PH_3; **M. W.**, 34; **G. F.**, $P\begin{cases} H \\ H \\ H. \end{cases}$

PROPERTIES AND PREPARATION.—This colorless gas, formed during the manufacture of hypophosphites, usually of disagreeable odor, is spontaneously inflammable, due to the presence of a compound, P_2H_4, as an impurity. It is slightly soluble in water, the solution being faintly alkaline. With hydrochloric acid it forms a compound, PH_4Cl (phosphorium chloride), analogous to ammonium chloride.

QUESTIONS.

When and by whom was phosphorus discovered?

How is phosphorus found in nature, and how is it obtained?

State the physical and chemical properties of phosphorus, and its allotropic modifications.

State antidotes and method of detecting phosphorus in case of poisoning.

State name and composition of the oxides of phosphorus.

What are the properties, mode of manufacture, and composition of phosphorous and hypophosphorous acids?

What precautions should be observed in prescribing hypophosphites?

State names and composition of phosphoric acids.

How is orthophosphoric acid made, what are its properties and tests, and how are its salts named?

How may orthophosphoric acid be distinguished from metaphosphoric and *pyrophosphoric* acids?

CHAPTER XII.

THE HALOGENS.

CLASS SIMILARITIES.—Fluorine, Chlorine, Bromine, and Iodine form a natural group, the *Halogens* (*salt-makers*), having many similarities. All ordinarily univalent, may show a valence of three, five, seven; with hydrogen form HF, HCl, HBr, and HI, colorless gases, and monobasic acids, soluble in water, combine directly with many metals to form fluorides, chlorides, bromides, and iodides; and because of their strong affinities, always occur in nature in combination.

The combining power varies with the atomic weight. Fluorine, with the lowest atomic weight, has greatest affinity, and iodine has the least affinity for other elements. They have, in the gaseous state, characteristic color, irritating odor, and marked bleaching and disinfecting properties.

FLUORINE.

S., F; **V.**, 1; **A. W.**, 19.

PHYSICAL AND CHEMICAL PROPERTIES.—This yellowish gas, of irritating and suffocating odor, has stronger affinities than any other element; supports combustion more vigorously than oxygen; ignites silicon, boron, and similar resistant substances, decomposes cold water into HF and oxygen, and sodium chloride into NaF with liberation of chlorine, ignites many organic compounds spontaneously, their hydrogen being burnt and carbon deposited, and combines *directly* with many metals, including gold and platinum, forming fluorides.

OCCURRENCE.—In nature, as fluorspar (calcium fluoride) and cryolite (sodium and aluminum fluoride), and in traces in many minerals, spring-waters, the enamel of the teeth, and bones of mammals.

PREPARATION.—From hydrofluoric acid, by electrolysis, with a strong current, in platinum vessels, kept at low temperature to prevent their too rapid corrosion.

COMPOUNDS OF FLUORINE.

Hydrofluoric Acid (Hydrogen Fluoride).

M. F., HF; **M. W.**, 20; **G. F.**, F—H.

PREPARATION.—By the action of sulphuric acid on calcium fluoride—*i. e.*,

$$CaF_2 + H_2SO_4 = 2HF + CaSO_4.$$

PROPERTIES.—As a gas dissolved in water the acid is very corrosive, and etches glass, whose silica it attacks; is a powerful antiseptic and

destroys certain objectionable ferments, occurring in yeast, without impeding the activity of that which converts sugar into alcohol. By mixing a small quantity of it with the fermenting mixture the quantity of alcoholic product is increased.

CHLORINE.

S., Cl; **V.**, 1; **A. W.**, 35.4.

OCCURRENCE.—In nature, chiefly as sodium chloride (common salt) dissolved in all natural water, varying from small amounts in ordinary spring-water to large amounts in sea-water, and solid deposits (rock salt). Other natural chlorides are those of potassium, magnesium, and calcium.

PROPERTIES.—Chlorine taken into the animal body as sodium chloride takes part in many physiological chemical changes.

This yellowish-green gas irritates the air-passages, is soluble in water two volumes to one at 10° C. (*aqua chlori*, *chlorine water*, U. S. P.). Under six atmospheres chlorine forms a greenish liquid.

CHEMICAL PROPERTIES.—It combines *directly* with all other elements except oxygen, nitrogen, and carbon, but with these *indirectly*, often with emission of light and explosion—*e. g.*, a mixture of chlorine and hydrogen is explosive, not when the mixture is kept cool and dark, but when ignited or exposed to sun-light. Chlorine decomposes water, ammonia, and hydrocarbons, combining with their hydrogen to form hydrochloric acid. This explains the unstableness of chlorine water. Chlorine deodorizes and bleaches either by combining directly with the material or by decomposing the water, the nascent oxygen then oxidizing the odorous or colored matter.

PREPARATION.—1. By electrolysis of sodium chloride. 2. By decomposing hydrochloric acid by oxidizing agents, most commonly manganese dioxide or potassium chlorate—*i. e.*,

$$MnO_2 + 4HCl = MnCl_2 + 2H_2O + 2Cl;$$

$$KClO_3 + 6HCl = KCl + 3H_2O + 6Cl.$$

COMPOUNDS OF CHLORINE AND HYDROGEN.

Hydrochloric Acid (Muriatic Acid).

M. F., HCl; **M. W.**, 36.4; **G. F.**, H—Cl.

PREPARATION.—By the action of sulphuric acid on sodium chloride—*i. e.*,

$$2NaCl + H_2SO_4 = Na_2SO_4 + 2HCl.$$

PHYSICAL PROPERTIES.—This colorless gas, of great affinity for

water, forming dense white clouds of liquid hydrochloric acid, when in contact with steam or moist air, is soluble in water (400 volumes in 1). This solution is also called hydrochloric acid.

VARIETIES.—*Hydrochloric acid*, U. S. P. (Sp. G. 1.163), contains 31.9 per cent. of pure HCl.

Dilute hydrochloric acid, U. S. P., contains 10 per cent. of HCl.

Qualitative Analysis of Hydrochloric Acid.

1. It forms insoluble chlorides in solutions of mercurous, lead, and silver salts. The silver chloride is insoluble in nitric acid, soluble in ammonia.
2. Chlorides treated with sulphuric acid and manganese dioxide evolve chlorine.

Nitrohydrochloric Acid, Aqua Regia (Nitromuriatic Acid).

NATURE.—A mixture of chloronitrous or chloronitric gas with water and chlorine, therefore **M. F.**, **M. W.**, and **G. F.** can not be stated.

PREPARATION.—By mixing concentrated nitric and hydrochloric acids.

CHEMICAL PROPERTIES.—It dissolves gold and platinum (which resist other acids), by the presence in the mixture of free chlorine and chloronitrous or chloronitric acid, which easily part with their chlorine—*i. e.*,

$$HNO_3 + 3HCl = 2H_2O + 2Cl + NOCl;$$

$$HNO_3 + 3HCl = 2H_2O + Cl + NOCl_2.$$

VARIETIES.—*Dilute nitromuriatic acid*, U. S. P., is composed of hydrochloric acid 180 c.c. and nitric acid 40 c.c. mixed. When the reaction is complete water 780 c.c. are added.

COMPOUNDS OF CHLORINE AND OXYGEN.

Oxides of Chlorine.

These form *indirectly:*

Chlorine monoxide,

$$\textbf{M. F.},\ Cl_2O;\ \textbf{M. W.},\ 87;\ \textbf{G. F.},\ <^{Cl}_{Cl}O.$$

Chlorine dioxide,

$$\textbf{M. F.},\ ClO_2;\ \textbf{M. W.},\ 67.5;\ \textbf{G. F.},\ \text{not determined.}$$

Chlorine trioxide,

$$\textbf{M. F.},\ Cl_2O_3;\ \textbf{M. W.},\ 119;\ \textbf{G. F.},\ \begin{array}{l} Cl-O \\ Cl-O \end{array}\!\!>\!O.$$

Chlorine heptoxide,

$$\textbf{M. F.},\ Cl_2O_7;\ \textbf{M. W.},\ 183.9;\ \textbf{G. F.},\ O=\begin{array}{c} O \\ \| \\ Cl \\ \| \\ O \end{array}-O-\begin{array}{c} O \\ \| \\ Cl \\ \| \\ O \end{array}=O.$$

Acids of Chlorine.

The first, third, and fourth oxides unite with water into:

Hypochlorous acid

$$(\textbf{M. F.},\ HClO;\ \textbf{M. W.},\ 52.5;\ \textbf{G. F.},\ Cl-O-H),$$

$$Cl_2O + H_2O = 2HClO;$$

Chlorous acid

$$(\textbf{M. F.},\ HClO_2;\ \textbf{M. W.},\ 120;\ \textbf{G. F.},\ O=Cl-O-H),$$

$$Cl_2O_3 + H_2O = 2HClO_2;$$

Perchloric acid

$$(\textbf{M. F.},\ HClO_4;\ \textbf{M. W.},\ 100.5;\ \textbf{G. F.},\ O\!<\!\begin{array}{c} O \\ O \end{array}\!>\!Cl-O-H),$$

$$Cl_2O_7 + H_2O = 2HClO_4;$$

An oxide yet unknown unites with water into:
Chloric acid,

$$\textbf{M. F.},\ HClO_3;\ \textbf{M. W.},\ 84.5;\ \textbf{G. F.},\ \begin{array}{c} O \\ | \\ O \end{array}\!>\!Cl-O-H.$$

Properties of the Oxides and Acids.—All are very unstable, frequently decomposing with explosion, univalent, and of little interest. The salts of hypochlorous and chloric acids only are of importance.

The Hypochlorites.

These are hypochlorous acid salts.

Preparation.—By the action of chlorine on the **hydroxides** of sodium, potassium, or calcium—*i. e.*,

$$2KOH + 2Cl = H_2O + KCl + KClO.$$

Chemical Properties.—These strong bleaching agents evolve chlorine on addition of acids—*i. e.*,

$$2KClO + H_2SO_4 = K_2SO_4 + H_2O + 2Cl.$$

The Chlorates.

These are chloric acid salts.

Preparation.—1. By the action of chlorine on alkali hydroxides at 100° C.—*i. e.*,

$$6KOH + 6Cl = 3H_2O + 5KCl + KClO_3.$$

2. By boiling a hypochlorite solution—*i. e.*,

$$3KClO = KCl + KClO_3.$$

Qualitative Analysis of Chlorates.

1. They liberate oxygen when heated. 2. They give off an explosive yellow gas (Cl_2O_4) when treated with sulphuric acid.

COMPOUNDS OF CHLORINE AND NITROGEN.

Nitrogen Trichloride.

M. F., NCl_3; **M. W.**, 120.5; **G. F.**, $N\begin{cases}Cl \\ Cl. \\ Cl\end{cases}$

Properties.—This oily liquid is a most explosive substance.

Preparation.—By the action of chlorine on ammonia—*i. e.*,

$$NH_3 + 6Cl = 3HCl + NCl_3.$$

BROMINE.

S., Br; **V.**, 1; **A. W.**, 79.7.

Occurrence.—In nature, chiefly as calcium, magnesium, and sodium bromides in mineral waters.

Physical Properties.—This heavy (Sp. G. 2.99), dark reddish-brown, volatile liquid, freezing at —240° C., and giving off suffocating yellowish-red fumes at ordinary temperatures, is soluble in thirty parts of water, more so in alcohol, and very freely in ether and bisulphide of carbon.

Preparation.—1. By decomposing native magnesium bromide, from evaporation of mineral water containing it, by chlorine—*i. e.*,

$$MgBr_2 + 2Cl = MgCl_2 + 2Br.$$

2. Concentrated sulphuric acid added to a bromide liberates bromine and sulphur dioxide—*i. e.*,

$$2NaBr + 2H_2SO_4 = Na_2SO_4 + 2Br + SO_2 + 2H_2O.$$

COMPOUNDS OF BROMINE AND HYDROGEN.

Hydrobromic Acid.

M. F., HBr; **M. W.**, 80.7; **G. F.**, Br—H.

PROPERTIES.—This colorless gas is soluble in water.

PREPARATION.—1. By the action of dilute sulphuric acid on bromides—*i. e.*,

$$MgBr_2 + H_2SO_4 = 2HBr + MgSO_4.$$

2. By the action of alcoholic tartaric acid solution or aqueous bromide solution—*i. e.*,

$$KBr + H_2C_4H_4O_6 = KHC_4H_4O_6 + HBr.$$

The mixture is kept at 0° C. until the potassium acid tartrate has separated and the solution of hydrobromic acid filtered off.

VARIETIES.—*Dilute hydrobromic acid*, U. S. P., contains 10 per cent. of pure HBr.

Qualitative Analysis of Hydrobromic Acid and Bromides.

1. Bromides give a yellow precipitate with silver nitrate, soluble in ammonia, insoluble in nitric acid.

2. Chlorine water added to a bromide solution containing a little starch-paste produces a yellow color.

Hypobromites and **bromates** are respectively salts of hypobromous ($HBrO$) and bromic acid ($HBrO_3$).

PREPARATION.—By processes analogous to those for obtaining the similar chlorine compounds.

IODINE.

S., I; **V.**, 1; **A. W.**, 126.8.

OCCURRENCE.—As the iodide or iodate of metals in spring- and sea-water and Chile salt-petre. From sea-water it is taken up by aquatic animals and plants, notably *kelp*, a species of sea-weed.

PREPARATION.—By burning the kelp, extracting the ashes with water, removing other salts by concentrating this solution and crystallizing, then liberating the iodine from the compounds in the mother-liquor by chlorine, and refining by sublimation.

PHYSICAL PROPERTIES.—This solid forms bluish-black scales, of

metallic lustre, Sp. G. 4.95, melting-point 115° C., boiling-point 200° C., is almost insoluble in water, except in the presence of potassium iodide. Its solution in alcohol or ether is brown, and in chloroform or carbon disulphide violet. Iodine and its solutions stain the skin brown, and taken internally are irritant poisons.

CHEMICAL PROPERTIES.—With the other halogens it forms unstable compounds, as iodine chloride, ICl.

COMPOUNDS OF IODINE AND HYDROGEN.

Hydriodic Acid (Hydrogen Iodide).

M. F., HI; **M. W.**, 127.8; **G. F.**, I—H.

PREPARATION.—1. By processes analogous to those for hydrobromic acid. 2. By the action of hydrosulphuric acid on iodine in the presence of water:

$$H_2S + 2I = 2HI + S.$$

Iodides are the salts of hydriodic acid.

Qualitative Analysis of Iodine and Iodides.

1. Starch-paste is colored blue by iodine.
2. If an iodide solution containing starch-paste be treated with chlorine-water, a blue color develops, by liberation of iodine and its action on the starch.

Iodic Acid.

M. F., HIO_3; **M. W.**, 175.8; **G. F.**, $\begin{matrix} O \searrow \\ O \nearrow \end{matrix}$ I—OH.

PROPERTIES.—This white crystalline substance, soluble in water, liberates iodine with reducing agents.

PREPARATION.—By dissolving in water iodine pentoxide, which is obtained by the action of nitric acid on iodine—*i. e.*,

$$6I + 10HNO_3 = 5N_2O_2 + 5H_2O + 3I_2O_5;$$
$$I_2O_5 + H_2O = 2HIO_3.$$

OTHER COMPOUNDS OF IODINE.

Sulphur Iodide.

M. F., S_2I_2; **M. W.**, 317.6; **G. F.**, $\begin{matrix} S—I \\ | \\ S—I \end{matrix}$.

PROPERTIES.—This grayish-black solid has metallic lustre.

PREPARATION.—By fusing a mixture of sulphur and iodine in proportions of their atomic weights.

Nitrogen Iodide.

M. F., NI_3; **M. W.**, 394.4; **G. F.**, $N\begin{cases} I \\ I \\ I. \end{cases}$

PREPARATION.—By the action of iodine on ammonia, it is a dark powder, exploding on concussion.

$$NH_3 + 6I = 3HI + NI_3.$$

QUESTIONS.

State the names and general properties of the halogens.

How does fluorine occur in nature, how is it obtained, and what are its properties?

State the composition and properties of hydrofluoric acid.

How does chlorine occur in nature, how is it obtained, and what are its properties?

State the composition, mode of manufacture, properties, and tests for the recognition of hydrochloric acid.

What is aqua regia, and what are its properties?

State the names and composition of the oxides of chlorine, and the acids formed from them.

How are hypochlorites obtained, what are their properties, and by what tests may they be recognized?

Give tests for chlorates.

Show by equation the action of chlorine on potassium hydroxide solution, at the ordinary temperature, and at the boiling-point.

How does bromine occur in nature, how is it obtained, and what are its properties?

State the composition, mode of manufacture, properties, and tests of hydrobromic acid.

Show by equation the action of concentrated sulphuric acid on potassium bromide.

Show by equation the formation of sodium hypobromite from bromine and sodium hydroxide.

How does iodine occur in nature, how is it obtained, and what are its properties?

How is hydriodic acid obtained, what are its properties, and by what tests may it be recognized?

State the composition and mode of manufacture of iodic acid.

State the composition and properties of sulphur iodide, nitrogen iodide, and nitrogen chloride.

CHAPTER XIII.

THE METALS, ALLOYS, AND AMALGAMS.

THE METALS.

Common Characteristics.

ABOUT one-half these fifty-eight elements are of practical importance, and the remainder, occurring in small quantity, of scientific interest. Their general properties have been mentioned (Page 72).

PHYSICAL PROPERTIES.—All are solid at ordinary temperatures (except mercury, a liquid). All are fusible, and some volatile.

CHEMICAL PROPERTIES.—All combine with fluorine, chlorine, and oxygen; some with sulphur, bromine, and iodine; a few with carbon and phosphorus, forming fluorides, chlorides, oxides, sulphides, bromides, iodides, carbides, and phosphides, respectively. Metals form salts by replacing the hydrogen in acids.

OCCURRENCE.—A few occur in nature in a free state; the majority, in combination as oxides, sulphides, carbonates, silicates, sulphates, and chlorides.

PREPARATION.—1. By heating the oxides with charcoal. The carbon combines with oxygen and the metal is liberated—*e. g.*,

$$FeO + C = CO + Fe.$$

2. The sulphides are converted into sulphates or oxides by *roasting*—*i. e.*, by heating to a high degree in an oxidizing atmosphere.

3. By electrolysis of their ores.

The Classification of Metals.

Groups may be variously made. Depending on their analytical peculiarities, metals are divided into two classes and six groups.

Class I. Light Metals.

Sp. G. 4 to 6. Sulphides soluble in water. This class consists of the following groups:

1. EARTH METALS.—*Aluminum* and *some rare metals*. Oxides insoluble.

2. ALKALINE EARTH METALS.—*Calcium*, *barium*, *strontium*, and *magnesium*. Oxides soluble; carbonates insoluble.

3. ALKALI METALS.—*Potassium*, *sodium*, *lithium*. All salts with mineral acids soluble.

Class II. Heavy Metals.

Sp. G. 6 to 21.5. Sulphides insoluble. This class consists of the following groups:

1. ARSENIC GROUP.—*Arsenic, antimony, tin, gold, platinum, molybdenum.* Sulphides insoluble in dilute acids, soluble in ammonium sulphide.
2. LEAD GROUP.—*Lead, copper, bismuth, silver, mercury, cadmium.* Sulphides insoluble in dilute acids and in ammonium sulphide.
3. IRON GROUP.—*Iron, manganese, zinc, chromium, cobalt, nickel.* Sulphides soluble in dilute acids.

ALLOYS AND AMALGAMS.

DEFINITION.—Alloys are combinations of two or more metals. Amalgams are alloys containing mercury.

PROPERTIES.—The exact condition of metals in alloys is not known. Probably they are true atomic combinations dissolved or mixed with an excess of one constituent. Certain changes in physical properties (fusing-point, brittleness, hardness, etc.) indicate some degree of change in the metals of an alloy. The retention of their general physical characters shows that changes are only slight.

QUESTIONS.

What are the characteristic properties of metals?
What is an alloy, and what is an amalgam?
How are metals obtained from their oxides?
How are metals classified according to their analytical behavior?
What are the characteristics of each group?

CHAPTER XIV.

THE ALKALI METALS AND AMMONIUM.

CHEMICAL PROPERTIES.—The alkali metals are univalent, decompose water at ordinary temperatures, liberating hydrogen, and forming hydroxides soluble in water and strongly alkaline; combine directly with oxygen and chlorine; form salts with common acids, soluble, and white (unless the acid colors), fusible at red heat, and giving to a non-luminous flame characteristic color. Their carbonates and hydroxides are not decomposed by heat. The metals are white, soft solids, which must be preserved in coal-oil because of their affinity for oxygen and action on water.

PREPARATION.—By electrolysis of the hydroxides, the metal being liberated with hydrogen at the negative pole.

LITHIUM.

S., Li; **A. W.**, 7; **V.**, 1.

OCCURRENCE.—In nature, as *silicate* in a few rare minerals, and as *chloride* in spring-water.

PROPERTIES.—Lithium is the *lightest* of the solid elements. Its bromide, carbonate, and citrate are used as medicine. Its salts are generally less soluble than those of the other alkali metals. A strong solution of a lithium salt forms a precipitate of lithium phosphate on the addition of sodium phosphate.

PREPARATION.—By electrolysis of the chloride.

SODIUM.

S., Na; **A. W.**, 23; **V.**, 1.

OCCURRENCE.—In nature, as chloride, in sea-water, and solid beds (rock-salt), in small quantity in spring-waters, and in vegetables and animals; also in deposits, as nitrate and silicate.

COMPOUNDS OF SODIUM.

SINGLE SALTS OF SODIUM.

Sodium Chloride.—

M. F., $NaCl$; **M. W.**, 58.5; **G. F.**, Na—Cl.

PROPERTIES.—This very important compound often contains magnesium and calcium chlorides, which render it hygroscopic. Pure sodium chloride is not hygroscopic. It is an essential of animal food, and assists in all parts of the body in many decompositions of vital importance; is soluble in 2.8 parts of cold water, and 2.5 of hot water; crystallizes in cubes, often arranged in pyramidal steps (*hopper-shaped*).

Sodium Carbonate (Sal Soda, Washing Soda).—

M. F., $Na_2CO_3.10H_2O$; **M. W.**, 280; **G. F.**, $CO{<}^{ONa}_{ONa}$.

PREPARATION.—*Manufactured* in larger quantity than any other compound of the alkali metals, for use in glass, soap, and other substances, it is *prepared* from sodium chloride, by the Leblanc or Solvay process.

The Leblanc Process.—Sodium sulphate ("*salt-cake*") is formed by decomposing the chloride with sulphuric acid, then mixed with coal and limestone, and fused in a reverberatory furnace, with the following reaction:

$$Na_2SO_4 + 4C + CaCO_3 = CaS + 4CO + Na_2CO_3.$$

The product ("*black-ash*" or "*ball-soda*"), containing unchanged carbon and lime, when mixed with water yields a solution of sodium carbonate and a precipitate of calcium oxysulphide. This solution on evaporation yields *crude sodium carbonate* ("*soda-ash*"), which, when redissolved and recrystallized, forms *the pure* salt, containing ten molecules of water of crystallization.

The Solvay or Ammonia Process.—Sodium chloride is decomposed by ammonium bicarbonate under pressure, forming sodium bicarbonate and ammonium chloride—*i. e.*,

$$NaCl + NH_4HCO_3 = NaHCO_3 + NH_4Cl.$$

The ammonium chloride is treated with lime to form ammonia, and this is converted into bicarbonate by the carbon dioxide evolved in the conversion of sodium bicarbonate into carbonate, by heat. *In this way the process is made continuous.*

PROPERTIES.—Strongly alkaline, and easily soluble in water, it forms monoclinic crystals, efflorescing on exposure, more rapidly on heating, into dried sodium carbonate (*sodii carbonas exsiccatus*, U. S. P.).

Sodium Bicarbonate (Baking Soda).—

M. F., $NaHCO_3$; **M. W.**, 84; **G. F.**, $CO{<}^{OH}_{ONa}$.

PREPARATION.—1. By the ammonia process from sodium chloride. 2. By passing carbon dioxide through a solution of sodium carbonate.

PROPERTIES.—This white crystalline powder, soluble in 12 parts of cold water, is slightly alkaline.

Sodium Hydroxide (Caustic Soda).—

M. F., NaOH; **M. W.**, 40; **G. F.**, Na—OH.

PREPARATION AND PROPERTIES.—By processes analogous to those for obtaining potassium hydroxide (see page 139), which it physically and chemically resembles.

Sodium Monoxide.—

M. F., Na_2O; **M. W.**, 62; **G. F.**, $^{Na}_{Na}{>}O$.

PREPARATION.—By oxidizing sodium at low temperatures.

Sodium Dioxide (Peroxide).—

M. F., Na_2O_2; **M. W.**, 78; **G. F.**, $\begin{array}{l} Na—O \\ \quad\;\;\, | \\ Na—O \end{array}$.

PREPARATION.—By oxidizing sodium at 200° C.

PROPERTIES.—This bleaching or oxidizing agent, dissolved in water, forms sodium hydroxide and oxygen, and in dilute acids hydrogen dioxide.

Sodium Sulphate (Glauber's Salt).—

M. F., Na_2SO_4; **M. W.**, 142; **G. F.**, $\begin{matrix}O\\O\end{matrix}\!\!>\!S\!<\!\begin{matrix}ONa\\ONa\end{matrix}$ or $SO_2<\begin{matrix}ONa\\ONa\end{matrix}$.

PREPARATION.—By the action of sulphuric acid on sodium chloride.

PROPERTIES.—Its large colorless efflorescent crystals dissolve in 2.8 parts of cold and 2.5 of hot water.

Sodium Bisulphite (Acid Sodium Sulphite).—

M. F., $NaHSO_3$; **M. W.**, 104; **G. F.**, $SO<\begin{matrix}OH\\ONa\end{matrix}$.

PREPARATION.—By the action of sulphur dioxide on cold solution of sodium carbonate:

$$Na_2CO_3 + 2SO_2 + H_2O = 2NaHSO_3 + CO_2.$$

Neutral Sodium Sulphite.—

M. F., $Na_2SO_3.7H_2O$; **M. W.**, 252; **G. F.**, $SO<\begin{matrix}ONa\\ONa\end{matrix}$.

PREPARATION.—By adding to the foregoing solution a quantity of sodium carbonate equal to that first used.

Sodium Thiosulphate (Sodium Hyposulphite).—

M. F., $Na_2S_2O_35H_2O$; **M. W.**, 248; **G. F.**, $SO_2<\begin{matrix}S-Na\\O-Na\end{matrix}$.

PREPARATION.—By digesting powdered sulphur with sodium sulphite solution.

Disodium-hydrogen Phosphate (Sodium Phosphate).—

M. F., $Na_2HPO_4 - 12H_2O$; **M. W.**, 258; **G. F.**, $O=P\!\begin{matrix}\diagup ONa\\ -ONa.\\ \diagdown OH\end{matrix}$

PREPARATION.—By the action of sodium carbonate on calcium acid phosphate, obtained by treating neutral calcium phosphate with sulphuric acid:

$$Ca_3(PO_4)_2 + 2H_2SO_4 = 2CaSO_4 + Ca(H_2PO_4)_2;$$

$$Ca(H_2PO_4)_2 + Na_2CO_3 = CaHPO_4 + H_2O + CO_2 + Na_2HPO_4.$$

REACTION.—Alkaline to litmus, acid to phenolphthalein.

Normal Sodium Phosphate.—

$$\textbf{M. F.},\ Na_3PO_4;\ \textbf{M. W.},\ 164;\ \textbf{G. F.},\ O{=}P\begin{cases}ONa\\ONa.\\ONa\end{cases}$$

PROPERTIES.—This unstable compound forms sodium-hydrogen phosphate and carbonate in the air.

Sodium Pyrophosphate.—

$$\textbf{M. F.},\ Na_4P_2O_7.10H_2O;\ \textbf{M. W.},\ 466;\ \textbf{G. F.},\ \begin{array}{c}\quad NaO\quad\ NaO\\ \quad\ |\qquad\ \ |\\ O{=}P{-}O{-}P{=}O.\\ \quad\ |\qquad\ \ |\\ \quad NaO\quad\ NaO\end{array}$$

PREPARATION.—By heating sodium phosphate to redness.

Sodium Nitrate (Chili Salt-petre).—

$$\textbf{M. F.},\ NaNO_3;\ \textbf{M. W.},\ 85;\ \textbf{G. F.},\ \begin{matrix}O\\O\end{matrix}\!\!>\!N{-}ONa.$$

OCCURRENCE.—In nature and purified by crystallization.

Sodium Nitrite.—

$$\textbf{M. F.},\ NaNO_2;\ \textbf{M. W.},\ 69;\ \textbf{G. F.},\ N\begin{cases}{=}O\\ONa\end{cases}$$

PREPARATION.—By fusing sodium nitrate and lead, and separating sodium nitrite from lead oxide so formed, by solution in water.

Sodium Hypophosphite, Iodide, Bromide, Chlorate, and Hypochlorite.—

PREPARATION.—By processes analogous to those for corresponding potassium compounds (Pages 141 and 142).

Labarraque's solution (eau de javelle) is a solution of sodium hypochlorite.

DOUBLE SALTS OF SODIUM.

Sodium-cobaltic Nitrite.—

M. F., $Co_2(NO_2)_6(NaNo_2)$; **M. W.**, 780.6; **G. F.**, not determined.

USES.—As a reagent for potassium and ammonium compounds.

Soda-lime *is* a mixture of lime and caustic soda.

POTASSIUM.

S., K; **V.**, 1; **A. W.**, 39.

OCCURRENCE.—In nature, as silicate of potassium and aluminum in granite and other rocks and, after their disintegration, in the soil; as chloride, nitrate, and sulphate in spring-water, sea-water, and the animal body; and as organic compounds in plants, for which it is an essential food.

SOURCES.—Chiefly sylvite, KCl, and carnallite ($KClMgCl_2 6H_2O$) in the Strassfurt mines.

COMPOUNDS OF POTASSIUM.

Potassium Hydroxide (Potassa, Caustic Potash).

M. F., KOH; **M. W.**, 56; **G. F.**, $K—O—H$.

PREPARATION.—1. By electrolysis of potassium chloride.
2. By the action of the metal on water:

$$K + H_2O = KOH + H.$$

3. By the action of calcium hydroxide on potassium carbonate:

$$K_2CO_3 + Ca(OH)_2 = CaCO_3 + 2KOH.$$

PROPERTIES.—This solid is brittle, white, very deliquescent, soluble in 5 parts of water, less soluble in alcohol, strongly alkaline; a powerful base, combining with all acids and destroying organic tissue; fuses into an oily liquid, which poured into moulds, forms pencils on cooling—the ordinary market form.

Potassa cum calce, U. S. P., is a mixture of equal quantities of potassium and calcium hydroxides.

Liquor potassæ, U. S. P., is a 5 per cent. watery solution of potassium hydroxide.

Oxides of Potassium.

Potassium monoxide:

M. F., K_2O; **M. W.**, 94; **G. F.**, $K—O—K$.

Potassium dioxide:

M. F., K_2O_2; **M. W.**, 110; **G. F.**, $\begin{array}{c} K—O \\ \quad\ | \\ K—O \end{array}$.

Potassium tetraoxide:

M. F., K_2O_4; **M. W.**, 142; **G. F.**, $\begin{array}{c} K—O—O \\ \qquad\quad | \\ K—O—O \end{array}$.

Potassium Carbonate.

M. F., K_2CO_3; **M. W.**, 138; **G. F.**, $CO<^{OK}_{OK}$.

PROPERTIES.—This strongly alkaline solid is soluble in equal weights of water.

PREPARATION.—1. From ashes of plants. 2. By the Leblanc process from native chloride (see page 135). Its organic compounds in the plant are converted into carbonate by burning. By extracting these ashes with water a solution of potassium carbonate, with traces of potassium and sodium chlorides and sulphates, is obtained, which evaporated to dryness forms "*pearl ash*" or "*crude potash.*"

Potassium Bicarbonate.

M. F., $KHCO_3$; **M. W.**, 100; **G. F.**, $CO<^{OH}_{OK}$.

PREPARATION.—This crystalline substance is obtained by treating potassium carbonate solution with carbon dioxide:

$$K_2CO_3 + H_2O + CO_2 = 2KHCO_3.$$

Potassium Nitrate (Nitre, Salt-Petre).

M. F., KNO_3; **M. W.**, 101; **G. F.**, $\begin{matrix}O\\ O\end{matrix}\!\gg\! N-OK.$

OCCURRENCE.—In nature with sodium nitrate in the soil, probably formed by nitrification of ammonia, absorbed from the air.

PROPERTIES.—It crystallizes in six-sided prisms, is soluble in water, has a cooling saline taste, oxidizes, and is therefore used in the manufacture of gunpowder.

Potassium Sulphate.—

M. F., K_2SO_4; **M. W.**, 174; **G. F.**, $SO_2<^{OK}_{OK}$.

PREPARATION.—By the action of sulphuric acid on potassium carbonate.

Potassium Bisulphate (Potassium Acid Sulphate).—

M. F., $KHSO_4$; **M. W.**, 136; **G. F.**, $SO_2<^{OH}_{OK}$.

PREPARATION.—By the action of sulphuric acid on potassium *chloride.*

Potassium Sulphite.—

M. F., K_2SO_3; **M. W.**, 158; **G. F.**, $SO{<}^{OK}_{OK}$.

PREPARATION.—By decomposing K_2CO_3 with sulphurous acid:

$$K_2CO_3 + H_2SO_3 = K_2SO_3 + H_2O + CO_2.$$

Potassa Sulphurata (Liver of Sulphur, Hepar Sulphuris).—

PREPARATION.—By fusing one part of sulphur and two parts of potassium carbonate.

PROPERTIES.—This mixture of potassium sulphide (K_2S), polysulphide (K_2S_3), and thiosulphate ($K_2S_2O_3$), absorbs water and oxygen, and becomes sulphate in the air.

Potassium Hypophosphite.—

M. F., KPH_2O_2; **M. W.**, 104; **G. F.**, $\begin{matrix} H \\ H \end{matrix}{>}P{<}\begin{matrix} {/\!\!/}O \\ OK \end{matrix}$.

PREPARATION.—By decomposing calcium hypophosphite by potassium carbonate. It is deliquescent, and easily soluble in water.

Potassium Chloride.—

M. F., KCl; **M. W.**, 74.5; **G. F.**, $K—Cl$.

OCCURRENCE.—In nature, and is used to form other potassium compounds and fertilizers.

Potassium Chlorate.

M. F., $KClO_3$; **M. W.**, 122.5; **G. F.**, $\begin{matrix} O \\ O \end{matrix}{\gg}Cl—OK$.

PREPARATION.—1. By the action of chlorine on boiling potassium hydroxide solution. 2. By electrolysis of potassium chloride solution under certain conditions.

PROPERTIES.—This white crystalline solid, strong oxidizing agent, yields oxygen when heated alone, decomposes when heated with organic matter, and with strong acids yields chloric acid, an explosive gas.

Hypochlorites and Chlorates.—

PREPARATION.—(See pages 128 and 129.)

Potassium Iodide.

M. F., KI; **M. W.**, 165.8; **G. F.**, K—I.

PREPARATION.—By adding iodine to potassium hydroxide solution until the brown color is permanent, evaporating to dryness, fusing the residue with charcoal to reduce the iodate formed, dissolving and crystallizing:

$$6KOH + 6I = 5KI + KIO_3 + 3H_2O;$$
$$KIO_3 + 3C = KI + 3CO.$$

PROPERTIES.—Its colorless cubical crystals are soluble in less than their own weight of water.

Potassium Bromide.—

M. F., KBr; **M. W.**, 118.7; **G. F.**, K—Br.

PREPARATION.—1. By a process analogous to the above. 2. By decomposing ferrous bromide by potassium carbonate:

$$FeBr_2 + K_2CO_3 = 2KBr + FeCO_3.$$

RUBIDIUM.

S., Rb; **V.**, 1; **A. W.**, 85.

CÆSIUM.

S., Cs; **V.**, 1; **A. W.**, 132.7.

Both these elements are of scientific interest only.

AMMONIUM.

M. F., NH_4; **M. W.**, 18; **G. F.**, $N \begin{matrix} \diagup H \\ < H \\ < H \\ \diagdown H \end{matrix}$.

PROPERTIES.—Ammonium salts show close analogy to sodium and potassium salts, and are therefore considered with the alkali metal salts.

The radical $\overline{NH_4}$ acts in combination like a univalent element; like all radicals it has no separate existence.

Ammonium amalgam, soft, spongy, and with metallic lustre, is obtained by decomposing ammonium chloride by potassium amalgam.

SOURCE.—The ammoniacal liquor of gas-works, a solution of ammonium hydroxide (NH_4OH).

All ammonium compounds volatilize at a low red heat.

COMPOUNDS OF AMMONIUM.

Ammonium Chloride (Sal-ammoniac).

M. F., $\overline{NH_4}Cl$; **M. W.,** 53.5; **G. F.,** $\overline{NH_4}-Cl$.

PREPARATION.—By neutralizing ammoniacal liquor with hydrochloric acid, evaporating, and purifying the residue by crystallization.

PROPERTIES.—It forms long, fibrous, tough, flexible, anhydrous crystals or a white granular powder, has sharp saline taste, neutral reaction, and is easily soluble in water.

Ammonium Sulphate, Nitrate, and Phosphate.

	AMMONIUM.		
	Sulphate.	Nitrate.	Phosphate.
Molecular Formula . .	$(NH_4)_2SO_4$.	NH_4NO_3.	$(NH_4)_2HPO_4$.
Molecular Weight . . .	132.	80.	132.
Graphic Formula . . .	$\begin{matrix}O\\O\end{matrix}\!\!>S<\!\!\begin{matrix}ONH_4\\ONH_4\end{matrix}$	$\begin{matrix}O\\O\end{matrix}\!\!>N-ONH_4$.	$O=P\begin{matrix}\diagup ONH_4\\-ONH_4\\\diagdown OH\end{matrix}$

PREPARATION.—By adding their respective acids to ammoniacal liquor or ammonium carbonate.

Microcosmic Salt (Phosphorous Salt).—Is used in blow-pipe analysis.

M. F., $NaNH_4HPO_4.4H_2O$; **M. W.,** 209; **G. F.,** $O=P\begin{matrix}\diagup ONa\\-ONH_4\\\diagdown OH\end{matrix}$.

Ammonium Carbamate.—

M. F., $NH_4CO_2NH_2$; **M. W.,** 78; **G. F.,** $CO<\begin{matrix}NH_2\\ONH_4\end{matrix}$.

PREPARATION.—It is formed when carbon dioxide and dry ammonia gas are brought together:

$$CO_2 + 2NH_3 = NH_4NH_2CO_2.$$

Carbamic acid is *carbonic acid* in which hydroxyl has been replaced by NH_2:

$$CO(OH)_2 + NH_3 = CONH_2OH + H_2O.$$

Commercial Ammonium Carbonate (Sal Volatile).—

M. F., $NH_4HCO_3.NH_4NH_2CO_2$; **M. W.,** 157;

G. F., $CO<\begin{matrix}NH_4\\OH\end{matrix}.CO<\begin{matrix}NH_2\\ONH_4\end{matrix}$.

PREPARATION.—This mixture of ammonium acid carbonate and carbamate is formed by sublimation of a mixture of ammonium chloride and calcium carbonate.

PROPERTIES.—Its white, translucent crystals on exposure to air become ammonium bicarbonate, a white powder; dissolved in water the carbamate becomes *normal* carbonate. Sufficient ammonia added converts the entire compound into normal carbonate:

$$NH_4HCO_3.NH_4NH_2CO_2 + H_2O + NH_3 = 2(NH_4)_2CO_3.$$

Aromatic spirit of ammonia, U. S. P., is a solution of normal ammonium carbonate in diluted alcohol containing some essential oils.

Ammonium-hydrogen Sulphide.—

M. F., NH_4HS; **M. W.**, 51; **G. F.**, $NH_4S{-}H$.

PREPARATION.—By the action of hydrogen sulphide on ammonium hydroxide:

$$NH_4OH + H_2S = NH_4HS + H_2O.$$

Ammonium Sulphide.—

M. F., $(NH_4)_2S$; **M. W.**, 68; **G. F.**, $NH_4{-}S{-}NH_4$.

PREPARATION.—By the action of ammonia gas on dry H_2S:

$$2NH_3 + H_2S = (NH_4)_2S.$$

Both substances are used as reagents.

Qualitative Analysis of the Alkali Metals and Ammonium.

This group is distinguished from others by forming soluble carbonates and phosphates. Their individual reactions are shown in the table below:

	Lithium.	Sodium.	Potassium.	Ammonium.
Sodium-cobaltic nitrite	No change	No change	Yellow precipitate	Yellow precipitate
Platinic chloride	No change	No change	Yellow precipitate	Yellow precipitate
Sodium phosphate	White precipitate	No change	No change	No change
Sodium hydroxide	No change	No change	No change	Ammonia gas
Heating to red heat	Fuses	Fuses	Fuses	Volatilizes
Flame color	Crimson	Yellow	Violet	

For the effect of ammonium compounds on Nessler's solution see page 93.

QUESTIONS.

State names, symbols, and valence of the alkali metals.

What are the general properties of the alkali metals?

State occurrence in nature, mode of manufacture, and composition of important salts of lithium.

State occurrence in nature, and mode of manufacture of sodium.
State composition and properties of sodium chloride.
How is sodium carbonate obtained?
Describe the Leblanc and Solvay processes.
State the composition, mode of manufacture, and properties of Glauber's salt.
What is the composition of sodium phosphate?
What is the composition of Chile salt-petre?
What is Labarraque's solution?
How does potassium occur in nature?
Give three processes for the manufacture of potassium hydroxide.
What is crude potash, and how is it obtained?
What is salt-petre?
What is liver of sulphur?
How is potassium iodide manufactured?
What is ammonium, and how does it resemble the alkali metals?
What is the source of ammonium compounds?
State the composition, mode of manufacture, and properties of ammonium chloride.
What is the composition of ammonium carbamate?
Show by equations the result of dissolving commercial ammonium carbonate in water, and in water containing ammonia.
How may the compounds of the alkali metals and ammonium be distinguished from those of other groups of elements?
How may potassium compounds be distinguished from those of sodium and ammonium?
How may lithium compounds be recognized?

CHAPTER XV.

MAGNESIUM AND COMPOUNDS OF MAGNESIUM.

MAGNESIUM.

S., Mg; **V.**, 2; **A. W.**, 24.3.

ANALYTICAL PROPERTIES.—These are intermediate between those of the alkali metals and alkaline earths, resemble the latter in the insolubility of the carbonate and phosphate in water, but differ in the carbonate being soluble in ammonium chloride solution. Magnesium salts form a crystalline precipitate with phosphoric acid in the presence of ammonium chloride and ammonia. Several of its salts are isomeric with analogous salts of zinc, which it otherwise resembles.

OCCURRENCE.—In nature, as magnesite ($MgCO_3$); kieserite ($MgSO_4H_2O$); camallite ($KClMgCl_2 6H_2O$); serpentine, talc, meerschaum, and other silicates; dolomite (calcium and magnesium carbonate); as soluble salts in sea and spring waters; as carbonate, and phosphates in plants and animals.

CHEMICAL PROPERTIES.—This silver-white metal easily oxidizes in moist air, losing its lustre, decomposes hot water, evolving hydrogen, burns in air or oxygen with a brilliant white light ("*flash-light*" of photography), forming magnesium oxide.

PREPARATION.—By electrolysis of the double chloride of magnesium and potassium.

COMPOUNDS OF MAGNESIUM.

Magnesium Sulphate (Epsom Salt).

M. F., $MgSO_4.7H_2O$; **M. W.**, 246.3; **G. F.**, $SO_2{<}{}^{O}_{O}{>}Mg$.

OCCURRENCE.—In nature in sea and spring waters and as kieserite.

PREPARATION.—As colorless crystals soluble in water by the action of sulphuric acid on native carbonate.

$$MgCO_3 + H_2SO_4 = MgSO_4 + H_2O + CO_2$$

Magnesium Carbonate (Light Magnesia, Magnesia Alba, U. S. P.).

M. F., (approximate) $(MgCO_3)_4Mg(OH)_2.5H_2O$; **M. W.** and **G. F.**, not known.

PREPARATION.—By boiling a solution of magnesium sulphate with sodium carbonate:

$$5MgSO_4 + 5Na_2CO_3 + 6H_2O = (MgCO_3)_4Mg(OH)_2.5H_2O + 5Na_2SO_4 + CO_2.$$

The precipitate is collected on a filter, washed, and dried to a light white powder (*light magnesia*, U. S. P.). If the above mixture be evaporated to dryness and sodium sulphate removed by washing, the residue is more compact (*heavy magnesium carbonate*, U. S. P.).

Magnesium Oxide (Calcined Magnesia).

M. F., MgO; **M. W.**, 40.3; **G. F.**, $Mg{=}O$.

PREPARATION.—By heating magnesium carbonate to full red heat:

$$(MgCO_3)_4Mg(OH)_2.5H_2O = 5MgO + 6H_2O + 4CO_2.$$

According as light or heavy carbonate is used, light or heavy magnesia (*magnesia ponderosa*, U. S. P.) results.

PROPERTIES.—This white amorphous powder absorbs water, and carbon dioxide from the air. With water it forms hydroxide—$Mg(OH)_2$.

Milk of magnesia, U. S. P., is a suspension of $Mg(OH)_2$ in water.

Magnesium Phosphates.

These resemble calcium phosphates in properties and composition.

Magnesium Ammonium Phosphate (Triple Phosphate).

M. F., $MgNH_4PO_4$; **M. W.**, 137.3; **G. F.**, $O{=}P\begin{cases} ONH_4 \\ OMg \\ OMg \end{cases}$.

OCCURRENCE.—It is deposited from ammoniacal urine, and recognized by coffin-lid-shaped crystals, and in the bladder forms phosphatic calculi.

PREPARATION.—See reactions for phosphoric acid, page 122.

QUESTIONS.

What elements does magnesium resemble, and how?

How is magnesium found in nature, how is it obtained, and what are its properties?

State occurrence in nature, composition, mode of manufacture, and properties of Epsom salt.

How is magnesium carbonate made?

How may light and heavy magnesia be obtained from magnesium sulphate?

Give tests for magnesium compounds.

CHAPTER XVI.

THE ALKALINE EARTHS.

GROUP CHARACTERISTICS.—Beryllium, Calcium, Strontium, Barium, form the alkaline earth group, are all bivalent light metals, occur in nature in several combinations, and are prepared by electrolysis and other decompositions, but always with difficulty.

They *resemble* the alkali metals, in decomposing water at ordinary temperatures, liberating hydrogen and forming soluble hydroxides, but *differ* from them in forming insoluble carbonates and phosphates. Barium sulphate is not soluble, and calcium and strontium sulphates are sparingly so. From the earths they differ in forming soluble hydroxides, whose solubility is greatest with barium (1–28.6), least with calcium (1–666), and intermediate (1–50) with strontium.

BERYLLIUM.

S., Be; **V.**, 2; **A. W.**, 9.

This is of scientific interest only.

CALCIUM.

S., Ca; **V.**, 2; **A. W.**, 40.

OCCURRENCE.—In nature, abundantly as carbonate (limestone, chalk, marble), sulphate (gypsum and alabaster), phosphate (apatite), fluoride (fluorspar), and silicate; as carbonate in shells of eggs and mollusca, phosphate and sulphate in animal tissues (bones), and chloride in water, blood, and other animal fluids.

COMPOUNDS OF CALCIUM.

Calcium Oxide (Lime, Burnt Lime, Quicklime).

M. F., CaO; **M. W.**, 56; **G. F.**, Ca=O.

PREPARATION.—By heating natural carbonate to about 800° C. (lime-burning):

$$CaCO_3 = CaO + CO_2.$$

PROPERTIES.—This amorphous, white, infusible, caustic, alkaline powder absorbs water and carbon dioxide from the air, forming *air-slacked lime*, unites with water, evolving great heat, into the *hydroxide*, $Ca(OH)_2$, which is soluble 15 parts in 10,000 of water, but more soluble in sugar solution—*syrupus calcis*, U. S. P.

Commercially lime is the most important base, being used in many chemical manufactures.

Lime water (*aqua calcis*, U. S. P.) is a saturated solution of calcium hydroxide, $Ca(OH)_2$, in water.

Test for Lime Water.—Blow expired air (CO_2) through a tube into lime water (see page 107).

Milk of lime is a mixture of calcium hydroxide and water.

Mortar is a mixture of lime, sand, and hair. On standing, a hard, tenacious calcium silicate is formed.

Hydraulic cement contains lime, clay and other silicates. After moistening, silicate of lime and aluminum is formed, causing the hardening.

Calcium Chloride.

M. F., $CaCl_2$; **M. W.**, 75.5; **G. F.**, $Ca<^{Cl}_{Cl}$.

PREPARATION.—By dissolving natural carbonate in hydrochloric acid and crystallizing:

$$CaCO_3 + 2HCl = CaCl_2 + H_2O + CO_2$$

PROPERTIES.—These colorless deliquescent crystals containing six molecules of water of crystallization, on heating become *anhydrous chloride*, a porous mass, which absorbs water gradually, and is therefore used for drying gases.

Calcium Bromide.

M. F., $CaBr_2$; **M. W.**, 119.7; **G. F.**, $Ca{<}^{B}_{Br}$.

PREPARATION.—Similarly to the chloride.

Chlorinated Lime, Calx Chlorata (Bleaching Powder).

PROPERTIES.—This white, hygroscopic, strong bleaching and disinfecting powder, is a mixture of calcium chloride and hypochlorite, with faint odor of chlorine. Chlorine is evolved from it when it is moistened, or treated with acids.

PREPARATION.—By treating calcium hydroxide with chlorine:

$$2Ca(OH)_2 + 4Cl = 2H_2O + CaCl_2 + Ca(ClO)_2.$$

Its activity is due to the hypochlorite, of which enough should be present to yield 35 per cent. of *available chlorine* with acid:

$$Ca(ClO)_2 + 4HCl = CaCl_2 + 2H_2O + 4Cl.$$

Calcium Carbonate, U. S. P.

M. F., $CaCO_3$; **M. W.**, 100; **G. F.**, $CO{<}^{O}_{O}{>}Ca$.

Calcii Carbonas Præcipitatus, U. S. P., is pure carbonate obtained as a white, neutral, amorphous powder, by mixing solutions of calcium chloride and sodium carbonate, washing, and drying the precipitate:

$$CaCl_2 + Na_2CO_3 = 2NaCl + CaCO_3.$$

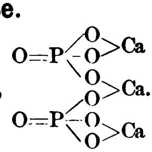
Calcium Sulphate (Dried Gypsum, Plaster of Paris).

M F., $CaSO_4$; **M. W.**, 136; **G. F.**, $SO_2{<}^{O}_{O}{>}Ca$.

PREPARATION.—By heating natural sulphate ($CaSO_4.2H_2O$) at 115° C. until nearly all water has been removed, forming a soft white powder, which after combining with water hardens on drying. It is therefore used in making moulds, casts, and immobilizing dressings in surgery.

Tricalcium Phosphate.

M. F., $Ca_3(PO_4)_2$; **M. W.**, 310; **G. F.**,



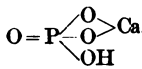
OCCURRENCE.—In nature in rocks and bones.

PROPERTIES.—This white, tasteless, amorphous powder insoluble in water, is soluble in hydrochloric and nitric acids.

PREPARATION.—Pure as *calcii phosphatum præcipitatum*, U. S. P., by treating calcium chloride with sodium phosphate and ammonia:

$$3CaCl_2 + 2Na_2HPO_4 + 2NH_4OH = Ca_3(PO_4)_2 + 4NaCl + 2NH_4Cl + 2H_2O.$$

Calcium Acid Phosphate (Superphosphate).

M. F., $CaHPO_4$; **M. W.**, 136; **G. F.**, 

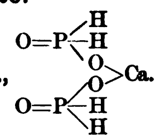
PROPERTIES.—A mixture of calcium acid phosphate and sulphate and used as a fertilizer.

PREPARATION.—By action of sulphuric acid on natural phosphate.

Bone ash is calcium phosphate and other salts, remaining after burning away organic matter of bone.

Bone black (animal charcoal) is obtained by destructive distillation of bone.

Calcium Hypophosphite.

M. F., $Ca(PH_2O_2)_2$; **M. W.**, 170; **G. F.**, 

PROPERTIES.—This white crystalline powder is soluble in water (1–6).

PREPARATION.—By boiling phosphorus with milk of lime until hydrogen phosphide is no longer evolved, removing excess of lime with carbon dioxide, and evaporating the clear filtrate.

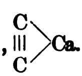
Calcium Phosphide.

M. F., Ca_2P_2; **M. W.**, 142; **G. F.**, unknown.

PREPARATION.—By heating a mixture of lime and phosphorus to redness. Its structural formula is unknown, and its formation can not be shown by equation.

PROPERTIES.—This chocolate-brown powder evolves hydrogen phosphide when thrown into water.

Calcium Carbide.

M. F., C_2Ca; **M. W.**, 92; **G. F.**, 

PREPARATION.—By heating lime and carbon in an electric furnace.

The elements combine into dark-brown or black crystalline masses, which with water form acetylene (C_2H_2), a gas burning with an intensely white flame:

$$C_2Ca + 2H_2O = C_2H_2 + Ca(OH)_2.$$

The manufacture of an organic substance by direct union of the elements is of interest and importance.

Calx Sulphurata (Sulphurated Lime).

M. F., M. W., and **G. F.,** can not be stated, because it is a mixture of calcium sulphide and sulphate, and is not constant in its proportions of each.

PREPARATION.—By heating anhydrous gypsum, starch, and charcoal in a crucible until colorless:

$$CaSO_4 + 2C = CaS + 2CO_2 \text{ (approximate).}$$

STRONTIUM.

S., Sr; **V.,** 2; **A. W.,** 87.

OCCURRENCE.—In nature as strontianite ($SrCO_3$) and celestite ($SrSO_4$), and in spring-water.

PROPERTIES.—The carbonate resists heat more than calcium carbonate. The oxide is obtained by heating the nitrate or hydroxide.

Strontium Nitrate.

M. F., $Sr(NO_3)_2$; **M. W.,** 211.3; **G. F.,** $Sr{<}^{ONO_2}_{ONO_2}$.

USES.—In the manufacture of pyrotechnics (*fireworks*) because of the brilliant red color of the flame of burning strontium.

PREPARATION.—By dissolving the carbonate in nitric acid:

$$SrCO_3 + 2HNO_3 = Sr(NO_3)_2 + H_2O + CO_2.$$

BARIUM.

S., Ba; **V.,** 2; **A. W.,** 136.9.

OCCURRENCE.—In nature as barite (heavy spar) ($BaSO_4$) or witherite ($BaCO_3$).

COMPOUNDS OF BARIUM.

Barium Chloride and Nitrate.

M. F., $BaCl_2.2H_2O$, $Ba(NO_3)_2$; **M. W.**, 243.9, 260.9;

G. F., $Ba{<}^{Cl}_{Cl}$, $Ba{<}^{ONO_2}_{ONO_2}$; *respectively.*

USES.—Valuable reagents.

PREPARATION.—By dissolving the natural carbonate in the respective acids and evaporating to crystallization:

$$BaCO_3 + 2HCl = BaCl_2 + H_2O + CO_2;$$

$$BaCO_3 + 2HNO = Ba(NO_3)_2 + H_2O + CO_2.$$

Barium Oxide.

M. F., BaO; **M. W.**, 152.9; **G. F.**, $Ba{=}O$.

PREPARATION.—By heating the nitrate:

$$Ba(NO_3)_2 = BaO + 2NO_2 + O.$$

Barium Dioxide.

M. F., BaO_2; **M. W.**, 168.9; **G. F.**, $Ba{<}^{O}_{O}{\Big|}$.

PREPARATION.—By heating the oxide to dull redness in air or oxygen:

$$BaO + O = BaO_2.$$

PROPERTIES.—Too great heat decomposes it into oxide and oxygen. This power of absorbing oxygen from air and liberating it at high temperatures is used for manufacturing oxygen. After being heated several times, barium oxide loses this absorbing power, and therefore cannot be used indefinitely.

Toxicology of Barium.

The salts of barium are all poisonous.

CHEMICAL ANTIDOTES.—Soluble sulphates precipitate barium sulphate (*very insoluble*) in the intestinal tract.

Qualitative Analysis of the Alkaline Earths and Magnesium.

TABLE OF REACTIONS DISTINGUISHING THE MEMBERS OF THE GROUP FROM EACH OTHER AND FROM MAGNESIUM.

	GROUP MEMBERS.			
Reagent.	Magnesium.[1]	Calcium.	Strontium.	Barium.
Ammonium carbonate.	White precipitate.	White precipitate.	White precipitate.	White precipitate.
Ammonium hydroxide.	White precipitate.			
Ammonium oxalate.	......	White precipitate in all solutions.	White precipitate in strong solutions.	White precipitate in strong solutions.
Sodium phosphate.	White precipitate.	White precipitate.	White precipitate.	White precipitate.
Potassium chromate.	......	......	Yellow precipitate.	Yellow precipitate.
Potassium dichromate.	......	......	....	Yellow precipitate.
Sulphuric acid.	......	White precipitate in concentrated solutions.	White precipitate in concentrated solutions.	White precipitate in dilute solutions.
Flame.	......	Yellowish red.	Red.	Yellowish green.

[1] Magnesium carbonate is soluble in ammonium chloride solution.

QUESTIONS.

Give names, symbols, and valences of the alkaline earth metals.

How do these metals differ from the alkali and the earth metals?

How does calcium occur in nature?

What is lime, and how is it obtained?

What are lime-water and milk of lime?

State composition, mode of manufacture, properties, and uses of calcium chloride and anhydrous calcium chloride.

Show by equations the manufacture of bleaching-powder from lime.

What are the properties and uses of chlorinated lime?

Show by equation the action of hydrochloric acid on calcium hypochlorite.

How much bleaching-powder of official strength will be required in the formation of 354 grams of chlorine?

What is plaster of Paris, and how is it made?

Show by equations the formation of precipitated carbonate and precipitated phosphate of calcium.

What is meant by superphosphate, bone-ash, and bone-black, and how are these substances obtained?

State composition, properties, and mode of manufacture of calcium hypophosphite.

What is calcium carbide, how is it obtained, and how does it act when treated with water?

How does strontium occur in nature?

How is strontium oxide manufactured?

Name the official salts of strontium, and state the method of their preparation.

How does barium occur in nature?
What are the antidotes used in cases of poisoning by barium? Why?
How is barium oxide obtained?
State composition, mode of manufacture, and uses of barium dioxide.
Give tests for distinguishing salts of calcium, strontium, barium, and magnesium from each other.

CHAPTER XVII.

THE EARTH METALS.

ALUMINUM and Cerium comprise this group, with several rare metals of scientific interest only.

ALUMINUM.

S., Al; **V.**, 3; **A. W.**, 27.

OCCURRENCE.—In nature *abundantly* as silicate in various rocks, and in the clays of their disintegration, as *traces* in mineral waters, plants, and animals. Ruby, sapphire, corundum, and emery are crystallized aluminum oxide colored by traces of other substances.

PREPARATION.—By electrolysis of a mixture of native fluoride of sodium and aluminum (Na_3AlF_6, cryolite) and aluminum oxide.

PHYSICAL PROPERTIES.—This very light (Sp. G. 2.67) tin-white metal, capable of high polish, malleable and ductile, conducts heat and electricity well, and is not greatly changed by dry or moist air.

USES.—In the pure state for manufacturing ornaments, and in alloys, which it greatly strengthens.

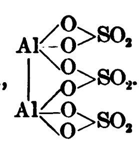
Aluminum bronze, an alloy of 90 parts of copper and 10 of aluminum, resembles gold in appearance and is much used in the arts.

CHEMICAL PROPERTIES.—Trivalent, under certain conditions the double atom Al_2 seems to combine as a single sexivalent atom—*e. g.*, the chloride, $AlCl_3$ or Al_2Cl_6. To explain this behavior, aluminum is looked on as a quadrivalent element, the two atoms being joined by one bond, as shown in the graphic formula below.

COMPOUNDS OF ALUMINUM.

Aluminum Sulphate.

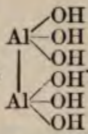
M. F., $Al_2(SO_4)_3.16H_2O$; **M. W.**, 630; **G. F.**,



PROPERTIES.—These pearl-white plates are soluble in about their *own weight* of water.

PREPARATION.—By treating clay or other native aluminum salts with sulphuric acid and crystallizing.

Aluminum Hydroxide.

M. F., $Al_2(OH)_6$; **M. W.**, 156; **G. F.**, 

PREPARATION.—By adding to solutions of aluminum salts soluble hydroxides, carbonates, or sulphides, respectively, thus:

1. $Al_2(SO_4)_3 + 6NaOH = Al_2(OH)_6 + 3Na_2SO_4$.
2. $Al_2(SO_4)_3 + 3Na_2CO_3 + 3H_2O = Al_2(OH)_6 + 3Na_2SO_4 + 3CO_2$.
3. $Al_2(SO_4)_3 + 3(NH_4)_2S + 6H_2O = Al_2(OH)_6 + 3(NH_4)_2SO_4 + 3H_2S$.

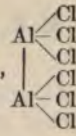
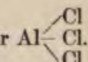
CHEMICAL PROPERTIES.—The usual reaction between solutions of the carbonate of an alkali metal and the soluble salt of any other metal, is the formation of an *insoluble carbonate*. *Aluminum oxide* is so weak a base that it cannot combine with carbonic acid, also weak; the reactions in the foregoing first two equations therefore resulting in the hydroxide and liberation of carbon dioxide; and the sulphides precipitate the hydroxide with liberation of hydrogen sulphide. Chromium and iron (ferric salts) show similar reactions.

Aluminum hydroxide precipitates as gelatinous flakes, which may be dried at 40° C. Greater heat expels water, forming aluminum oxide (alumina), Al_2O_3.

USES.—By surface action takes up vegetable colors and is used in dyeing, particularly of cotton, as a mordant to "fix" colors, being either precipitated in a solution of the dye, forming the pigments called *lakes*, or soaked into the material, which is then placed in the color.

The hydroxide is precipitated in water in order to purify it, by removing suspended matter and some soluble constituents.

Aluminium Chloride.

M. F., Al_2Cl_6 or $AlCl_3$; **M. W.**, 266.4 or 133.2; **G. F.**,  or 

PREPARATION.—As an anhydrous solid.

1. By the action of chlorine in a heated mixture of carbon and alumina:

$$Al_2O_3 + 3C + 6Cl = Al_2Cl_6 + 3CO.$$

2. In solution by adding hydrochloric acid to the oxide or hydroxide. When the water is removed by heat, the chloride formed in this way decomposes into oxide and hydrochloric acid:

$$Al_2Cl_6 + 3H_2O = Al_2O_3 + 6HCl.$$

Alums.

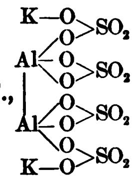
Definition.—These *isomorphous* salts consist of *one* molecule of the sulphate of a *univalent* metal, *one* of the sulphate of a *trivalent* metal, and *twenty-four* of water of crystallization.

General Formula.—

$$M_2SO_4M_2'''(SO_4)_3.24H_2O.$$

Official Alum (Potassium and Aluminum Sulphate, Potassium Alum).

M. F., $K_2SO_4.Al_2(SO_4)_3.24H_2O$; **M. W.**, 948; **G. F.**,



$+ 24H_2O.$

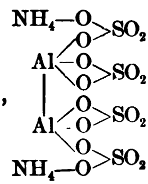
Occurrence.—In large octahedrons, easily soluble in water; the solution is acid.

Preparation.—By treating clay with sulphuric acid, adding to the solution of aluminum sulphate so formed potassium sulphate, evaporating and crystallizing.

Ammonium-aluminum Alum (Ammonia Alum).—

M. F., $(NH_4)_2SO_4Al_2(SO_4)_3.24H_2O$; **M. W.**, 906;

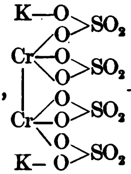
G. F.,



$+ 24H_2O.$

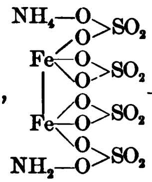
Potassium-chromium Alum (Chrome Alum).—

M. F., $K_2SO_4.Cr_2(SO_4)_3.24H_2O$; **M. W.,** 850;

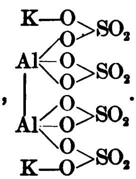
G. F.,  $+ 24H_2O$.

Ammonium-ferric Alum (Ferric Alum).—

M. F., $(NH_4)_2SO_4Fe_2(SO_4)_3.24H_2O$; **M. W.,** 815.8;

G. F.,  $+ 24H_2O$.

Dried Alum (Alumen Exsiccatum, Burnt Alum, U. S. P.).—

M. F., $K_2SO_4(Al)_2SO_4$; **M. W.,** 324; **G. F.,** .

PREPARATION.—As a white caustic powder from alum by expelling the water of crystallization by heat.

Clays.

COMPOSITION.—Clays are plastic masses consisting largely of aluminum silicate combined with water. *Kaolin* is almost pure aluminum silicate.

Fire clay contains an excess of silica.

Red clay owes its color to the contained ferric oxide.

PROPERTIES.—By heating, clay loses its water and becomes hard and porous, as in porcelain, stoneware, earthenware, bricks, etc.

Mica (Isinglass).

Mica is a native double silicate of aluminum and certain other elements.

Ultramarine Blue.

OCCURRENCE.—This beautiful pigment is found in nature as "*lapis lazuli.*"

PREPARATION.—By fusing a mixture of clay, sodium sulphate and carbonate, sulphur, and charcoal.

CHEMICAL CONSTITUTION is unknown.

CERIUM.

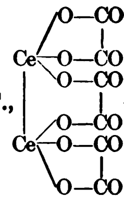
S., Ce; **V.**, 3; **A. W.**, 141.

OCCURRENCE.—In nature in small quantities, principally as cerite (cerium silicate).

CHEMICAL PROPERTIES resemble those of aluminum.

COMPOUNDS OF CERIUM.

Cerium Oxalate.

M. F., $Ce_2(C_2O_4)_3.10H_2O$; **M. W.**, 726; **G. F.**,  $+ 10H_2O$.

PREPARATION.—Precipitated from cerium solutions by ammonium oxalate as a white granular powder. This is the only cerium salt official.

ZIRCONIUM, ERBIUM, AND THORIUM.

These, also, are earth metals.

OCCURRENCE.—In small quantities in Monazite sand.

Oxide of Thorium is used in the mantle of the Welsbach burner.

Qualitative Analysis of the Earth Metals.

They are distinguished from the other groups by forming **insoluble** hydroxides, and by liberating carbon dioxide when **precipitated with** *soluble carbonates.* Aluminum is distinguished from cerium **by forming**

a precipitate with caustic potash, soluble in excess of the reagent. Cerium hydroxide is insoluble in potassium hydroxide.

QUESTIONS.

How does aluminum occur in nature, how is it obtained, and what are its properties?

What is the effect of the addition of sodium carbonate and ammonium sulphide to solutions of an aluminum salt?

What is aluminum oxide used for?

How is aluminum chloride obtained, and what is the effect of heat on its solution?

What are alums, and state the composition of four important alums?

State the composition and mode of manufacture of cerium oxalate.

CHAPTER XVIII.

THE IRON GROUP.

GROUP PROPERTIES.—Iron, Manganese, Chromium, Cobalt, Nickel, and Zinc are the metals of this group, and are distinguished by forming with soluble sulphides a precipitate soluble in dilute mineral acids. The precipitate is the sulphide of each metal except chromium, which, like aluminum, forms insoluble hydroxide when sulphides are added to its solution. As the sulphides are soluble in dilute acids, neutral or acid solutions of the metals are not precipitated by hydrogen sulphide, but by ammonium sulphide. All their carbonates, phosphates, oxides, and hydroxides are insoluble.

Except zinc, which is always bivalent, all form two series of compounds, in one as bivalents, in the other, trivalents, or a double atom acts as a single sexivalent atom. They unite with oxygen in several proportions. Some higher oxides show decided acid properties. All (*except zinc*) are magnetic, and at red heat decompose water into metallic oxides and hydrogen. All dissolve in dilute sulphuric and hydrochloric acids, liberating hydrogen and forming salts.

FERRUM (IRON).

S., Fe; **V.**, 2; **A. W.**, 55.9.

The most useful of the heavy metals.

OCCURRENCE.—In nature in *large* quantity as the most common inorganic coloring agent (the red or yellowish-red color of rocks, clay, sand, etc., is due to it); in *small* quantity in vegetables, and the blood and tissues of animals; also in metallic form as meteorites, but chiefly as loadstone, magnatite (ferrous-ferric oxide, $FeOFe_2O_3$), haematite

(ferric oxide, Fe_2O_3), sidenite (ferrous carbonate, $FeCO_3$), and iron pyrites (ferrous sulphide, FeS).

PREPARATION.—The ore is converted into ferric oxide, if necessary, by roasting, then mixed with charcoal (carbon) and limestone (calcium carbonate and silicate), and heated in a blast-furnace. At glowing heat the oxide is reduced by the carbon:

$$Fe_2O_3 + 3C = 2Fe + 3CO.$$

Being surrounded by the fused calcium silicate (*slag*), the molten iron is protected from oxidation by the air blast of the furnace. The iron and slag collect at the furnace-bottom and separate by gravity. The iron is run out into moulds to form *pig-iron.* The slag, drawn off later, is used in cement manufacture.

Pig or *cast iron* is impure, containing among other substances 2 to 5 per cent. of carbon.

Puddling is agitating molten iron in an oxidizing flame. The carbon is burnt out and pig iron becomes wrought iron.

Wrought or *bar iron* is produced by puddling.

Steel contains 0.16 to 2 per cent. of carbon. and is made by agitating a mixture of molten wrought and pig iron in proper proportions, and blowing air under pressure into the molten mass.

Pure Iron (*Ferrum Reductum*, U. S. P.).—This fine, grayish-black powder is obtained by reducing red-hot iron oxide by hydrogen.

COMPOUNDS OF IRON.

Iron forms two series of compounds:

1. FERROUS, as a *bivalent* atom.
2. FERRIC, as a *sexivalent* double atom. Ferrous compounds are converted into ferric in the air, or more readily by oxidizing agents, as chlorine, nitric acid, etc.

Ferrous Oxide.

M. F., FeO; **M. W.**, 71.9; **G. F.**, Fe=O.

OCCURRENCE.—Seldom free, as it absorbs oxygen forming ferric oxide.

PROPERTIES.—This strong base unites with acids into ferrous salts (usually pale green).

Ferrous Hydroxide.

M. F., $Fe(OH)_2$; **M. W.**, 89.9; **G. F.**, $Fe{<}^{OH}_{OH}$.

PREPARATION.—On the addition of alkaline hydroxides to ferrous *solutions this* white precipitate falls and rapidly absorbs oxygen and

water, becoming bluish-green, gray, black, and finally brown by formation of ferric hydroxide:

$$FeSO_4 + 2NaOH = Na_2SO_4 + Fe(OH)_2;$$

$$2Fe(OH)_2 + O + H_2O = Fe_2(OH)_6.$$

Ferric Oxide.

$$\textbf{M. F.},\ Fe_2O_3;\ \textbf{M. W.},\ 159.8;\ \textbf{G. F.},\ \begin{array}{l} Fe{=}O \\ \,|\quad\ \rangle O \\ Fe{=}O \end{array}$$

PREPARATION.—As a reddish-brown powder by expelling water from ferric hydroxide by heat.

PROPERTIES.—A feeble base uniting with acids into ferric salts, usually of brown color, resembles aluminum oxide in not combining with carbonic or hydrosulphuric acid. Ferric compounds are precipitated as hydroxides by soluble carbonates and are reduced to the ferrous state by sulphides.

Ferric Hydroxide (Hydrated Oxide of Iron, Per or Sesqui Oxide, Red Oxide of Iron).

$$\textbf{M. F.},\ Fe_2(OH)_6;\ \textbf{M. W.},\ 213.8;\ \textbf{G. F.},\ \begin{array}{r} HO\diagdown \\ HO-\!\!\!- \\ HO\diagup \end{array} Fe-Fe \begin{array}{l} \diagup OH \\ -\!\!\!- OH. \\ \diagdown OH \end{array}$$

PREPARATION.—As reddish-brown powder, by precipitating solutions of ferric salts by soluble carbonates or hydroxides:

$$Fe_2(Cl)_6 + 3Na_2CO_3 + 3H_2O = Fe_2(OH)_6 + 6NaCl + 3CO_2;$$

$$Fe_2(SO_4)_3 + 3Mg(OH)_2 = Fe_2(OH)_6 + 3MgSO_4.$$

USES.—An antidote in arsenic-poisoning, and should be freshly precipitated and washed, but not dried, because the moist salt combines more readily with arsenic than the dried.

Ferric Hydrate and Magnesia.—PREPARATION.—As in the foregoing equation. The entire mixture is the antidote. It is rapidly prepared, as no washing is required, the magnesium sulphate adding to its value as an antidote.

Ferrates.

These are compounds of alkalies with iron trioxide (FeO_3) and water.

Potassium Ferrate.—Free trioxide is not known.

$$\textbf{M. F.},\ K_2FeO_4;\ \textbf{M. W.},\ 197.9;\ \textbf{G. F.},\ FeO_2 < \begin{array}{l} O-K \\ O-K. \end{array}$$

Ferrous Chloride (Protochloride of Iron).

M. F., $FeCl_2$; **M. W.**, 126.9; **G. F.**, $Fe<^{Cl}_{Cl}$.

PREPARATION.—1. In the dry condition by heating iron in a current of dry hydrochloric acid gas. 2. As a pale-green solution by dissolving iron in hydrochloric acid. The salt and the solution absorb oxygen from the air, hence the salt cannot be obtained by evaporating the solution.

Ferric Chloride (Chloride, Sesqui or Per Chloride of Iron).

M. F., $Fe_2Cl_6.12H_2O$; **M. W.**, 540.8; **G. F.**, $\begin{matrix}Cl\diagdown & & \diagup Cl\\ Cl-Fe & - & Fe-Cl\\ Cl\diagup & & \diagdown Cl\end{matrix}$.

PROPERTIES.—These orange-yellow, deliquescent crystals have acid reaction, styptic taste, and by heat lose water of crystallization, decomposing into ferric oxide and hydrochloric acid.

PREPARATION.—By oxidizing solutions of ferrous chloride containing free hydrochloric acid with nitric acid. The hydrochloric acid is oxidized by the nitric acid, forming chlorine, which combines with the ferrous chloride, water and nitrogen dioxide, which are liberated:

$$3Fe_2Cl_6 + 2HNO_3 + 6HCl = 3Fe_2Cl_6 + 4H_2O + 2NO_2.$$

Solution of Chloride of Iron (**Liquor Ferri Chloridi,** U. S. P.).—PROPERTIES.—Contains 37.8 per cent. pure Fe_2Cl_6 with some free hydrochloric acid. Sp. G. 14.05.

Tinctura Ferri Chloridi, U. S. P.—PREPARATION.—This is a mixture of one part of the solution with three parts of alcohol, left standing in a closed vessel for three months; during this time the alcohol and ferric chloride form ferrous chloride and certain ethereal compounds.

Dialyzed Iron (Ferrum Dialyzatum, U. S. P.).—PREPARATION.—As an aqueous solution (about 5 per cent.) of oxychloride of iron by dialyzing a mixture of ferric chloride and ammonium hydroxide until all ammonium chloride is removed.

Ferrous Iodide.

M. F., FeI_2; **M. W.**, 308.9; **G. F.**, $Fe<^{I}_{I}$.

PREPARATION.—By direct combination in the presence of water, as green solution, and evaporation to dryness to form the salt, which is hygroscopic and easily decomposed by the air.

Saccharated Ferrous Iodide, U. S. P.—To prevent decomposition of freshly prepared iodide 20 per cent. of sugar of milk is added.

Ferrous Bromide.—

M. F., $FeBr_2$; **M. W.,** 215.4; **G. F.,** $Fe<^{Br}_{Br}$.

PREPARATION.—Analogous to the foregoing.

Ferrous Sulphide.

M. F., FeS; **M. W.,** 87.9; **G. F.,** Fe=S.

PREPARATION.—By fusing sulphur and iron in molecular proportions. Also found in nature.

USE.—For the manufacture of hydrogen sulphide.

When soluble sulphides are added to solutions of ferrous salts a black ferrous sulphide precipitates; when added to ferric solutions, the ferric salt is reduced to form ferrous sulphide, and sulphur is also precipitated:

$$Fe_2Cl_6 + 3Na_2S = 2FeS + 6NaCl + S.$$

Ferrous Sulphate (Sulphate of Iron, Green Vitriol, Copperas).

M. F., $FeSO_4.7H_2O$; **M. W.,** 277.9; **G. F.,** $SO_2<^{O}_{O}>Fe$.

PROPERTIES.—These large bluish-green efflorescent crystals, which absorb oxygen from the air, are soluble in water.

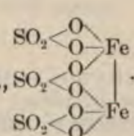
PREPARATION.—1. By dissolving iron in sulphuric acid. 2. By roasting native sulphide. 3. As a by-product in many chemical decompositions:

$$Fe + H_2SO_4 = FeSO_4 + 2H.$$

Granulated Ferrous Sulphate, U. S. P., is precipitated when strong solutions slightly acidified by sulphuric acid are poured into alcohol.

Dried Ferrous Sulphate, U. S. P. ($FeSO_4.3H_2O$), is formed by heating the crystals to 100° C.

Ferric Sulphate.

M. F., $Fe_2(SO_4)_3$; **M. W.,** 399.8; **G. F.,** 

PREPARATION.—By the action of nitric acid on solution of ferrous sulphate and sulphuric acid:

$$6FeSO_4 + 3H_2SO_4 + 2HNO_3 = 3(Fe_2(SO_4)_3) + 2NO + 4H_2O.$$

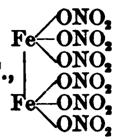
Solution of Ferric Sulphate (Liquor Ferri Tersulphatis), U. S. P.,

is a solution of this salt used in making ferric-ammonia alum and hydrated oxide of iron and magnesia.

Solution of Ferric Subsulphate (Monsel's Solution), U. S. P., similar to the above, contains less sulphuric acid, and is supposed to be a basic ferric sulphate solution.

COMPOSITION.—Probably $5Fe_2(SO_4)_3Fe_2(OH)_6$.

Ferric Nitrate.

M. F., $Fe(NO_3)_6$; **M. W.**, 427.9; **G. F.**, 

PREPARATION.—In solution by dissolving ferric hydroxide in nitric acid. A 6 per cent. solution is official.

Ferrous Carbonate.

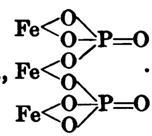
M. F., $FeCO_3$; **M. W.**, 115.9; **G. F.**, $CO<^{O}_{O}>Fe$.

PREPARATION.—As a nearly white, unstable, easily oxidized substance, by treating solutions of ferrous salts with alkaline carbonates.

Saccharated Carbonate of Iron, U. S. P. (15 per cent. ferrous carbonate).

PREPARATION.—By adding sugar to freshly precipitated carbonate and drying.

Ferrous Phosphate.

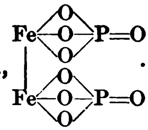
M. F., $Fe_3(PO_4)_2$; **M. W.**, 357.7; **G. F.**, .

PREPARATION.—As slate-colored powder darkening by absorption of oxygen, by adding a solution of sodium phosphate to a ferrous salt in the presence of sodium acetate:

$$3FeSO_4 + 2Na_2HPO_4 + 2NaC_2H_3O_2 = Fe_3(PO_4)_2 + 3Na_2SO_4 + 2HC_2H_3O_2.$$

Sodium acetate is added to prevent formation of free sulphuric acid, which would convert the ferrous phosphate into ferrous-hydrogen phosphate *($FeHPO_4$)*.

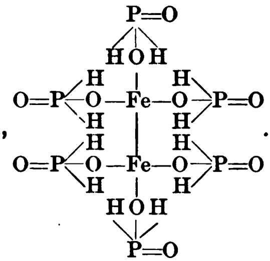
Ferric Phosphate.

M. F., $Fe_2(PO_4)_2$; **M. W.**, 301.8; **G. F.**, 

PREPARATION.—By precipitating ferric solutions with neutral phosphates.

Soluble Ferric Phosphate, U. S. P., is a scale compound (see Index).

Ferric Hypophosphite.

M. F., $Fe_2(H_2PO_2)_6$; **M. W.**, 501.8; **G. F.**, 

PREPARATION.—By dissolving ferric hydroxide in hypophosphorous acid and evaporating.

PROPERTIES.—A grayish-white powder slightly soluble in water, easily soluble in hydrochloric acid and in solutions of alkali citrates.

MANGANESE.

S., Mn; **V.**, 2; **A. W.**, 54.8.

OCCURRENCE.—In nature as oxide or carbonate.

PREPARATION.—By reducing native carbonate with carbon:

$$MnO_2 + C = Mn + CO_2;$$

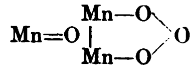
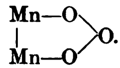
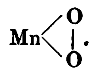
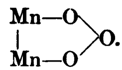
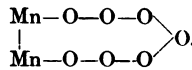
$$2Mn_2O_3 + 3C = 4Mn + 3CO_2.$$

PROPERTIES.—Resembles iron, is darker in color, harder, and more easily oxidized, and forms two series of salts, MANGANOUS as a *bivalent*, MANGANIC as a *sexivalent* double atom, also manganates (salts of manganic acid).

Recent researches show manganese to be an important constituent of oxidizing ferments in plants, and are of interest as possibly explaining the value of the metal as a medicine.

COMPOUNDS OF MANGANESE.

The Oxides of Manganese.

Common Name.	M. F.	M. W.	G. F.
1. Manganous oxide	MnO	70.8	Mn=O.
2. Mangano-manganic oxide . .	$MnO.Mn_2O_3 = Mn_3O_4$	228.4	
3. Manganic oxide	Mn_2O_3	157.6	
4. Manganese dioxide	MnO_2	86.8	
5. Manganese trioxide	MnO_3	102.8	
6. Manganese heptoxide	Mn_2O_7	221.6	

The trioxide and heptoxide are known only in combination as salts of manganic and permanganic acids.

Manganous Oxide.

PREPARATION.—1. By heating native carbonate or hydroxide. 2. By precipitating a solution of a manganous salt with sodium hydroxide.

PROPERTIES.—A strong base combining with acids into manganous salts, which usually are pink.

Manganese Dioxide (Manganese Peroxide, Binoxide, Black Oxide).

PROPERTIES.—This heavy, grayish-black, crystalline mineral, the most important manganese compound in nature, when heated to redness liberates oxygen, when mixed with potassium chlorate facilitates the evolution of oxygen (see page 89), and with hydrochloric acid liberates chlorine (see page 129).

***Mangani* Oxidum Nigrum, U. S. P., contains at least 66 per** *cent. of pure* manganese dioxide.

Manganous Sulphate.

M. F., $MnSO_4.4H_2O$; **M. W.**, 222.8; **G. F.**, $SO_2<{}^{O}_{O}>Mn$.

PROPERTIES.—These pale, rose-pink crystals combine with alkali sulphates into double salts.

PREPARATION.—By dissolving manganous oxide or dioxide in sulphuric acid and crystallizing.

Manganic Acid and the Manganates.

PREPARATION.—By fusing a manganese oxide with a mixture of alkali carbonates or hydroxides, and chlorate or nitrates. The manganese becomes manganic acid, which combines with the alkali into a manganate:

$$3MnO_2 + 6KOH + KClO_3 = 3K_2MnO_4 + 3H_2O + KCl.$$

The fused mass is extracted with water and evaporated to dark-green potassium manganate crystals.

Permanganic Acid.

M. F., $HMnO_4$; **M. W.**, 119.8; **G. F.**, MnO_3—OH.

PROPERTIES.—Stable, with acid properties, forms salts called permanganates.

PREPARATION.—By electrolysis of the potassium salt.

Potassium Permanganate.

M. F., $KMnO_4$; **M. W.**, 157.9; **G. F.**, MnO_3—OK.

PREPARATION.—1. By the action of dilute acid, or 2, of a large quantity of water on potassium manganate:

$$3K_2MnO_4 + 4HCl = MnO_2 + 4KCl + 2KMnO_4 + 2H_2O.$$

The dioxide is filtered out and the filtrate is evaporated to form slender, prismatic dark-purple crystals of somewhat metallic lustre. According to concentration, its solution varies from a faint pink to dark purple.

PROPERTIES.—A powerful oxidizing and disinfecting agent, easily giving up oxygen to substances having affinity for it, and losing its characteristic color by decomposition. In the presence of acids and reducing agents, two molecules liberate five atoms of oxygen:

$$2KMnO_4 + 3H_2O + x = K_2SO_4 + 2MnSO_4 + 3H_2O + xO_5$$

CHROMIUM.

S., Cr; **V.**, 2; **A. W.**, 52.

OCCURRENCE.—In nature chiefly as chromite (chromic iron ore), $FeOCr_2O_3$.

PREPARATION.—As a steel-gray metal with difficulty by reduction of the oxide by carbon:

$$FeOCr_2O_3 + 2C = Fe + 2Cr + 2CO_2.$$

PROPERTIES.—Resembles aluminum, iron, and sulphur. With soluble carbonates and sulphides solutions of its salts precipitate chromium hydroxide. It forms *two basic* oxides, CrO and Cr_2O_3, and *one acid* oxide, CrO_3. The last unites with water into chromic acid, H_2CrO_4, which, like sulphuric acid, forms insoluble salts with barium and lead.

COMPOUNDS OF CHROMIUM.

All are colored. The name is from *chroma*, (Gk.) *color*.

Compounds of chromic acid are more important in medicine than the others of chromium.

Potassium Dichromate (Bichromate or Red Chromate of Potassium).

M. F., $K_2Cr_2O_7$; **M. W.**, 294; **G. F.**, $C_2O_5<^{OK}_{OK}$.

PROPERTIES.—This source of other chromium compounds forms large red transparent crystals, soluble in water, is a strong oxidizer, liberates chlorine from hydrochloric acid, and, heated alone, evolves oxygen; in the presence of an acid and a reducing agent it liberates three atoms of oxygen from each molecule:

$$K_2Cr_2O_7 + 4H_2SO_4 + x = K_2SO_4 + Cr_2(SO_4)_3 + 4H_2O + xO_3.$$

Hydrogen sulphide reduces potassium dichromate, precipitating sulphur:

$$K_2Cr_2O_7 + 4H_2SO_4 + 3H_2S = K_2SO_4 + Cr_2(SO_4)_3 + 7H_2O + 3S.$$

PREPARATION.—By heating a mixture of chromium ore, calcium hydroxide, and potassium carbonate in an oxidizing flame, forming potassium chromate, ferric oxide, and carbon dioxide:

$$2(FeOCr_2O_3) + 4K_2CO_3 + 7O = Fe_2O_3 + 4CO_2 + 4K_2CrO_4.$$

The neutral chromate is dissolved out of the fused mass, forming a yellow solution, which is treated with sulphuric acid. The potassium *sulphate* and dichromate separate by crystallization.

Oxides of Chromium.

Chromium Trioxide (Chromic Acid, Chromic Anhydride).

M. F., CrO_3; **M. W.**, 100; **G. F.**, $Cr{<}^{O}_{O}{>}O$.

PREPARATION.—As deep-red, deliquescent, needle-shaped crystals when sulphuric acid is added to a concentrated solution of potassium dichromate.

PROPERTIES.—This strongly oxidizing, very corrosive agent forms with water chromic and dichromic acids, analogous to sulphuric and pyrosulphuric acids:

$$CrO_3 + H_2O = H_2CrO_4, \text{ Chromic acid;}$$

$$2CrO_3 + H_2O = H_2Cr_2O_7, \text{ Dichromic acid.}$$

Chromates and **Dichromates** are their respective salts.

Chromic Oxide.

M. F., Cr_2O_3; **M. W.**, 152; **G. F.**, $\begin{array}{l} Cr{<}^{O} \\ \quad\;\; {>}O. \\ Cr{<}_{O} \end{array}$

PREPARATION AND PROPERTIES.—As a green powder, insoluble in water, by the action of sulphur on dichromates. It is a *basic* oxide combining with acids into salts, and largely used to color glass and porcelain.

Chromic Hydroxide.

M. F., $Cr_2(OH)_6$; **M. W.**, 206; **G. F.**, $\begin{array}{l} Cr{\lt}\begin{smallmatrix}OH\\OH\\OH\end{smallmatrix} \\ \,| \\ Cr{\lt}\begin{smallmatrix}OH\\OH\\OH\end{smallmatrix} \end{array}.$

PREPARATION.—As a bluish-green gelatinous precipitate by adding ammonia water to solution of chromium salts. The formation of chromium sulphate by reducing potassium dichromate has been shown.

Chromium Salts.—PREPARATION.—By dissolving the hydroxide in the appropriate acid, evaporating and crystallizing. Solutions of chromium sulphate and potassium sulphate evaporate to form the purple crystals of chrome alum.

COBALT AND NICKEL.

S., Co, Ni; **V.**, 2, 2; **A. W.**, 58.6, 58.6, respectively.

PROPERTIES.—These resemble each other *chemically* and *physically*, occur in nature as arsenites and sulphides, frequently in the same ore, and are silver-white. Salts of cobalt are generally red, those of nickel green.

USES.—Metallic nickel is used in many alloys and electroplating. *German silver* is an alloy of nickel, copper, and zinc.

COMPOUNDS OF COBALT.

Cobaltous Chloride.

M. F., $CoCl_2$; **M. W.**, 129.6; **G. F.**, $Co{<}^{Cl}_{Cl}$.

PREPARATION.—By dissolving the oxide in hydrochloric acid and crystallizing. The solution and the hydrous crystals are rose-red. The anhydrous crystals are blue, and absorb water, becoming red. *Blue sympathetic ink* is a solution of cobaltous chloride in water. Writing with this ink is colorless until heated, when formation of anhydrous salts makes it blue and visible. *Green sympathetic ink* is a solution of cobaltous and nickelous chlorides.

ZINC.

S., Zn; **V.**, 2; **A. W.**, 65.3.

OCCURRENCE.—In nature as zinc blende (ZnS), calamine (zinc silicate), and Smithsonite ($ZnCO_3$).

PREPARATION.—By heating the carbonate or oxide with charcoal and distilling the liberated metal.

PROPERTIES.—A bluish-white, crystalline, brittle metal, slowly tarnishing in the air, being covered with a film of oxide and carbonate. Heated to 130° C., zinc becomes malleable after cooling. *Sheet zinc* is thus made. At 300° C. losing its malleability, it may be powdered; at 410° C. fuses; and at higher temperatures volatilizes, and in the air burns with a *greenish* light. It is bivalent, forming *one* series of usually colorless salts.

USES.—Metallic zinc in alloys; in galvanized iron as protective coating; and as a reagent in liberating hydrogen from acids.

COMPOUNDS OF ZINC.

Zinc Oxide (Flores Zinci, Zinc-White, U. S. P.).

M. F., ZnO; **M. W.**, 81; **G. F.**, Zn=O.

PROPERTIES.—A white amorphous tasteless powder, insoluble in

water, soluble in acids, when strongly heated turns yellow, but white again on cooling.

USES.—As a white pigment, and in medicine for external application.

PREPARATION.—By heating the carbonate:

$$ZnCO_3 = ZnO + CO_2.$$

Zinc Hydroxide.

M. F., $Zn(OH)_2$; **A. W.**, 99; **G. F.**, $Zn<{O-H \atop O-H}$.

PREPARATION.—Precipitates when alkali or ammonium hydroxide is added to zinc solutions, and is soluble in excess of the reagent.

Zinc Chloride (Butter of Zinc, U. S. P.).

M. F., $ZnCl_2$; **M. W.**, 136; **G. F.**, $Zn<{Cl \atop Cl}$.

PREPARATION.—By dissolving zinc or its carbonate in dilute hydrochloric acid, and evaporating to dryness.

PROPERTIES.—It is a deliquescent white powder or opaque mass, volatile, soluble in water and alcohol, and combines with proteins.

USES.—As a caustic, astringent, and preservative.

Liquor Zinci Chloridi, U. S. P., is a 50 per cent. aqueous solution of the salt.

Zinc Oxychloride.—PREPARATION.—By mixing from two to six parts of oxide with one of chloride, forming a plastic mass, which rapidly hardens, and is used in dentistry.

Zinc Oxyphosphate.—PREPARATION.—By mixing excess of zinc oxide with phosphoric acid; is similar to the oxychloride, and often replaces it in dentistry. The extent to which these mixtures form true chemical compounds is unknown. In both cases a great excess of zinc oxide is used.

Zinc Bromide.—

M. F., $ZnBr_2$; **M. W.**, 223.9; **G. F.**, $Zn<{Br \atop Br}$.

PREPARATION.—By dissolving zinc in hydrobromic acid. It resembles zinc chloride chemically and physically.

Zinc Iodide.—

M. F., ZnI_2: **M. W.**, 319; **G. F.**, $Zn<{I \atop I}$.

PREPARATION.—By direct combination of the elements in water. It resembles the chloride in general properties.

Zinc Carbonate (Zinci Carbonas Præcipitatum, U. S. P.).

M. F., $2(ZnCO_3).3(ZnOH_2)$; **M. W.,** and **G. F.,** can not be stated because the composition varies.

Preparation and Properties.—By precipitating zinc sulphate with sodium carbonate in boiling solution. The *basic* carbonate is a white powder, odorless, tasteless, soluble in acids and in ammonia water.

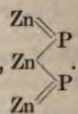
Zinc Sulphate (White Vitriol, U. S. P.).

M. F., $ZnSO_4.7H_2O$; **M. W.,** 287.1; **G. F.,** $SO_2\!<\!{}^{O}_{O}\!>\!Zn$.

Preparation.—As small white crystals, isomorphous with magnesium sulphate and soluble in water, by evaporating a solution of zinc in *dilute* sulphuric acid. For unexplained reasons, zinc is not acted on by *concentrated* sulphuric acid.

Uses.—As an astringent, stimulant, mild caustic, and emetic.

Zinc Phosphide.—

M. F., Zn_3P_2; **M. W.,** 257.3; **G. F.,** 

A grayish-black powder formed by direct combination when phosphorus is added to melted zinc.

TOXICOLOGY OF ZINC SALTS.

The soluble salts are poisonous, and usually cause vomiting.

Chemical Antidotes.—When vomiting is delayed, the stomach-tube should be used, and white of egg or tannic acid given to form respectively harmless albuminate and tannate of zinc.

Qualitative Analysis of the Iron Group.

This group is distinguished from the ALKALI and ALKALINE EARTH metals by precipitating with ammonia water and ammonium sulphide, usually as sulphide. This precipitate is soluble in dilute acids, which distinguishes the group from the lead and arsenic groups, both of which form sulphides insoluble in dilute acid. The earth metals resemble the iron metals analytically, therefore, the reactions of aluminum and cerium are included in the following summary. The iron and earth metals form TWO CLASSES by the solubility of the hydroxides of manganese, zinc, cobalt, and nickel, formed by adding ammonium hydroxide to *the respective* solutions in excess of the reagent and ammonium chlo-

Reagent.	Group Members.				
	Ferrous.	Ferric.	Chromium.	Aluminum.	Cerium.
1. Ammonium sulphide	Black precipitate	Black precipitate	Green precipitate	White precipitate	White precipitate
2. Ammonium hydroxide	Dirty green precipitate	Reddish-brown precipitate	Green precipitate	White precipitate	White precipitate
3. In excess of 2	Insoluble	Insoluble	Insoluble	Insoluble	Insoluble
4. Potassium hydroxide	Same as above	Reddish-brown precipitate	Green precipitate soluble in KOH, reprecipitated by boiling	White precipitate soluble in KOH	White precipitate insoluble in KOH
5. $(NH_4)_2CO_3$	White precipitate darkens on boiling	Same as above with evolution of CO_2	Green precipitate and CO_2	White precipitate and CO_2	White precipitate and CO_2
6. In excess of 4 and 5	Insoluble	Insoluble	Insoluble	Insoluble	Insoluble
7. $K_4Fe(CN)_6$	Pale-blue precipitate turning darker	Dark-blue precipitate			
8. $K_6Fe_2(CN)_{12}$	Dark-blue precipitate	Green-brown solution			
9. KCNS	. . . .	Dark red			

Reagent.	Group Members.			
	Manganese.	Zinc.	Cobalt.	Nickel.
1. Ammonium sulphide	Flesh-colored precipitate	White precipitate	Black precipitate	Black precipitate
2. Ammonium hydroxide	White precipitate	White precipitate	Blue precipitate	Green precipitate
3. In excess of 2	Soluble	Soluble	Soluble	Soluble
4. Potassium hydroxide	White precipitate	White precipitate	Blue precipitate	Green precipitate
5. In excess of 4	Insoluble	Soluble	Insoluble	Insoluble
6. Ammonium carbonate	White precipitate, darkens on standing	White precipitate	Blue precipitate	Green precipitate
7. In excess of 6	Insoluble	Soluble	Soluble	Soluble
8. $K_4Fe(CN)_6$	White precipitate	White precipitate	Grayish-green precipitate	Greenish-white precipitate
9. $K_6Fe_2(CN)_{12}$	Pale-brown precipitate	Pale brownish yellow precipitate	Deep brownish-red precipitate	Yellowish-brown precipitate
10. KCNS	. . . .	. . . .	Color intensified greatly	Color slightly intensified
11. Borax bead and fused with salt	Violet	. . . .	Blue	Red

ride. The hydroxides of iron, chromium, aluminum, and cerium are not soluble in ammonia. The two classes are summarized separately.

Chromic acid salts form characteristic precipitates with certain metals. *Yellow* with barium and lead salts, $BaCrO_4$ and $PbCrO_4$; *dark red* with silver, Ag_2CrO_4; *red* with mercurous salts, Hg_2CrO_4. Lead chromate is soluble in sodium hydroxide, but barium chromate is not.

QUESTIONS.

Which elements constitute the iron group of metals?

Which salts of these metals are soluble in water?

How do the sulphides of these metals differ from the sulphides of the lead and arsenic metals?

What is the valence of the different iron metals? State their general properties.

How does iron occur in nature, and how is it obtained from the ore?

What are meant by pig-iron, wrought-iron, and steel?

What is reduced iron, and how is it obtained?

How many oxides does iron form, and how do they differ?

How does ferric hydroxide differ from ferrous hydroxide, and how are they obtained?

What is ferric hydrate and magnesia used for? Show by equations its formation.

What are ferrates?

Show by equation the formation of ferric chloride from iron.

What is tinctura ferri chloridi?

What is dialyzed iron?

How is ferrous iodide obtained?

How is ferrous sulphide obtained?

How do soluble sulphides act on ferrous and ferric solutions?

State common name, composition, mode of manufacture, and properties of ferrous sulphate.

How is ferric sulphate made?

What is Monsel's solution?

How is ferrous carbonate manufactured?

What is the strength of saccharated ferrous carbonate and saccharated iodide of iron?

How does the presence or absence of sodium acetate affect the action of sodium phosphate on ferrous sulphate solution?

How does manganese occur in nature?

Give names and composition of the oxides of manganese existing in a separate state, and those existing only in combination.

State composition, properties, and uses of manganese dioxide.

How is manganous sulphate made, and what are its properties?

Show by equation the action of potassium permanganate and potassium dichromate in acid solution as oxidizing agents.

How many oxides of chromium are known?

State common name, composition, mode of manufacture, and uses of potassium dichromate.

How is chromic anhydride manufactured, and what are its properties?

How is chromic hydroxide obtained, and what are its uses?

How do cobalt and nickel occur in nature? What are the ordinary uses of these metals?

State the properties of zinc.

How is zinc oxide obtained, and what are its properties?
How are zinc phosphide and iodide made?
State composition, mode of manufacture, and properties of zinc chloride.
How are zinc oxychloride and oxyphosphate made, and what are their properties and uses?
State composition and mode of manufacture of zinci carbonas præcipitatum.
How do concentrated and dilute sulphuric acid affect zinc?
What are the properties and composition of zinc sulphate?
How may zinc, manganese, cobalt, and nickel be separated from the other iron and earth metals?
How are ferrous salts distinguished from ferric salts?
How is aluminum distinguished from cerium and chromium?
Give reactions for chromates.
How may zinc be distinguished from manganese analytically?
Give distinguishing tests of cobalt and nickel.

CHAPTER XIX.

THE LEAD GROUP.

LEAD, Copper, Bismuth, Mercury, Silver, and Cadmium comprise this group.

GROUP CHARACTERS.—When hydrogen sulphide is added to their solutions, they are precipitated as sulphides, insoluble in dilute acids, which distinguishes this from the iron group; and insoluble in ammonium sulphide, therein differing from arsenic metals. All these metals also form insoluble carbonates, phosphates, oxides, hydroxides, and iodides, and a few insoluble chlorides and sulphates. All their nitrates, most of their sulphates and chlorides are soluble, Their valences are not uniform. Several form two series of salts. These metals do not decompose water or *dilute* sulphuric acid; heated with *concentrated* sulphuric acid, form sulphates, usually liberating sulphur dioxide, and from nitric acid liberate nitrogen dioxide, forming nitrates.

LEAD.

S., Pb; **V.**, 2; **A. W.**, 206.4.

OCCURRENCE.—In nature chiefly as galena (PbS).

PREPARATION.—A mixture of oxide and sulphate with some unchanged sulphide is obtained by roasting the ore. On further heating, it decomposes to sulphur dioxide (SO_2) and metallic lead:

$$6PbS + 13O = PbSO_4 + 3PbO + 2PbS + 3SO_2;$$

$$PbSO_4 + 2PbO + 2PbS = 5Pb + 3SO_2.$$

PROPERTIES.—Lead is a soft, white metal, easily fusible, resistant to

acids, and used for pans and tubs for holding or conveying acids—*e. g.*, sulphuric acid chambers, and in several useful alloys, as type-metal, solder, etc.

The Toxicology of Lead.

The compounds of lead are directly poisonous. Lead-workers often suffer from *chronic* poisoning (*painters' colic*).

SOURCES OF POISONING.—A frequent source of lead-poisoning is water which has passed through lead pipes or been collected in leaden tanks. Lead slowly oxidizes in the air. By contact with air and water, lead hydroxide is formed, slightly soluble in water. When the water contains carbon dioxide, or soluble carbonates or sulphates, an insoluble deposit of carbonate or sulphate of lead is formed, and no lead enters into solution unless the carbon dioxide be under pressure, when some carbonate is dissolved as bicarbonate of lead. *Nitrates and nitrites in water increase its solvent action on lead, important facts, as drinking-water is often carried through lead pipes.*

TEST FOR LEAD IN WATER.—A dark color forms when H_2S gas is passed through the water, evaporated to $\frac{1}{10}$ or $\frac{1}{100}$ its volume.

CHEMICAL ANTIDOTES.—Sodium or magnesium sulphate.

PROPHYLAXIS.—A little sulphuric acid in water.

All three form insoluble lead sulphate, and owe their value as antidotes to this property.

COMPOUNDS OF LEAD.

Hardly any have great medicinal value.

Lead Oxide (Litharge).—A reddish-yellow crystalline powder.

M. F., PbO; **M. W.**, 222.4; **G. F.**, $Pb{=}O$.

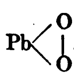
PREPARATION.—As powder by exposing melted lead to air-current, and fusing the yellow powder (*massicot*) so formed.

USED in making lead salts, lead plaster, and in the arts.

Red Lead.—PREPARATION.—As a red powder, by heating litharge in the air.

COMPOSITION.—Probably $PbO.Pb_2O$.

Lead Dioxide.—

M. F., PbO_2; **M. W.**, 238.4; **G. F.**, .

This dark-brown powder is the insoluble residue left when red lead is acted on by nitric acid.

Lead Nitrate.—

M. F., $Pb(NO_3)_2$; **M. W.**, 330.4; **G. F.**, $\begin{matrix} NO_2O \\ NO_2O \end{matrix}>Pb$.

The only soluble salt of lead and a mineral acid.

PREPARATION.—By dissolving the oxide in *dilute* nitric acid. Lead and lead nitrate are insoluble in *concentrated* nitric acid.

Lead Carbonate (White Lead).—

M. F., $2(PbCO_3)Pb(OH)_2$; **M. W.**, 773.2; **G. F.**, $CO<\begin{matrix} O \\ O \end{matrix}>Pb. \begin{matrix} HO \\ HO \end{matrix}>Pb$.

PREPARATION.—As an insoluble powder, (1) by precipitating a lead solution with a soluble carbonate, or (2) by the simultaneous action of air, carbon dioxide, and acetic acid on lead. It is frequently adulterated with barium sulphate.

Lead Iodide.—

M. F., PbI_2; **M. W.**, 460.4; **G. F.**, $Pb<\begin{matrix} I \\ I \end{matrix}$.

PREPARATION.—As a heavy, bright-yellow powder, by precipitating a lead solution by potassium iodide; is *distinguished* from lead chromate by being soluble in ammonium chloride.

COPPER.

S., Cu; **V.**, 2; **A. W.**, 63.2.

OCCURRENCE.—In nature free, and as cuprite (CuO), copper pyrites (Cu_2FeS_2), copper glance (CuS), and malachite (basic cupric carbonate).

PREPARATION.—By roasting the ores and reducing the oxide with coke. From acidified solutions copper is reduced by iron.

PROPERTIES.—Of red color, very malleable, it is one of the best conductors of heat and electricity, in dry air is unchanged, but in moist air gradually coats with green subcarbonate.

USES.—In many important alloys, brass, German silver, bronze, gun-metal, bell-metal, Babbitt-metal, and with gold and silver to add hardness to coins and ornaments. It is bivalent, forms two oxides, and two series of salts, *cuprous* and *cupric*.

COMPOUNDS OF COPPER.

Cupric Oxide (Black Oxide of Copper).

M. F., CuO; **M. W.**, 79.2; **G. F.**, Cu=O.

PREPARATION.—1. By heating copper to redness in the air; 2. By heating cupric carbonate or nitrate; and 3. By boiling the blue precipi-

tated cupric hydroxide ($Cu(OH)_2$) obtained by adding sodium or potassium hydroxide to copper solutions:

$$CuSO_4 + 2NaOH = Cu(OH)_2 + Na_2SO_4;$$

$$Cu(OH)_2 = CuO + H_2O.$$

Cuprous Oxide (Red Oxide or Suboxide of Copper).

M. F., Cu_2O; **M. W.**, 142.4; **G. F.**, $\begin{matrix}Cu\\ | \\ Cu\end{matrix}\!\!>O.$

PREPARATION.—As a red powder, by heating (1) cupric oxide with copper or with charcoal; (2) cupric oxide with organic matter; or (3) boiling cupric hydroxide in alkaline solution with certain reducing agents:

$$2CuO + C = Cu_2O + CO.$$

Cupric Sulphate (Blue Vitriol, Bluestone, U. S. P.).

M. F., $CuSO_4.5H_2O$; **M. W.**, 249.2; **G. F.**, $SO_2<^{O}_{O}>Cu.$

PROPERTIES.—This important compound forms large transparent blue crystals, easily soluble in water, having a nauseous taste, and at 200° C. losing their water of crystallization, thus forming anhydrous cupric sulphate, a white powder, used for desiccation.

PREPARATION.—1. By dissolving the oxide in sulphuric acid; 2. By roasting copper pyrites:

$$CuO + H_2SO_4 = CuSO_4 + H_2O.$$

Cupric Carbonate.—

M. F., $CuCO_3$; **M. W.**, 123.2; **G. F.**, $CO<^{O}_{O}>Cu.$

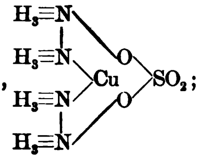
PREPARATION.—As a bluish-green powder, precipitated by adding sodium or potassium carbonate to solution of copper sulphate. Dissolved in acids it is converted into the corresponding *cupric* salts.

AMMONIO-COPPER COMPOUNDS.

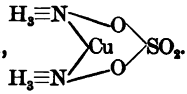
These are compounds of metallammoniums, hypothetical radicals, derived by substitution of hydrogen of ammonia by metals.

When ammonium hydroxide is added to a cupric sulphate solution, the pale-blue precipitate dissolves in excess of the reagent, the solution *turning deep blue* by the formation of:

1. Cupric Tetrammonium Sulphate.—

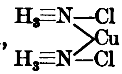
M. F., $CuSO_4(NH_3)_4$; **M. W.**, 227.2; **G. F.**,  ; or,

2. Cupric Diammonium Sulphate.—

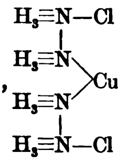
M. F., $CuSO_4(NH_3)_2$; **M. W.**, 193.2; **G. F.**, 

Similarly cupric chloride ($CuCl_2$) and ammonium form:

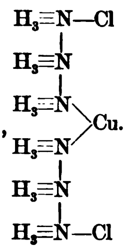
1. Cupric Diammonium Chloride.—

M. F., $CuCl_2(NH_3)_2$; **M. W.**, 168.2; **G. F.**,  ;

2. Cupric Tetrammonium Chloride.—

M. F., $CuCl_2(NH_3)_4$; **M. W.**, 202.2; **G. F.**,  ;

3. Cupric Hexammonium Chloride.—

M. F., $CuCl_2(NH_3)_6$; **M. W.**, 236.2; **G. F.**, 

The Toxicology of Copper.

The salts of copper are poisonous.

Chemical Antidotes.—After emptying the stomach, white of egg, reduced iron, and potassium ferrocyanide.

SOURCES OF POISONING.—Sometimes used to color food. It usually enters the food from *oxidized* cooking utensils. Perfectly clean, bright copper is not acted on by foods.

BISMUTH.

S., Bi; **V.**, 3 or 5; **A. W.**, 209.

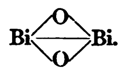
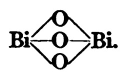
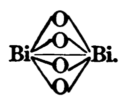
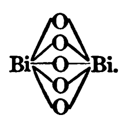
OCCURRENCE.—In nature as sulphides and oxides, but chiefly as the metal, and frequently mixed with arsenic.

PROPERTIES.—The metal is used in alloys and in salts, is extremely diamagnetic, slightly oxidized by the air, volatilizes at high temperatures, and is trivalent or quinquivalent.

COMPOUNDS OF BISMUTH.

Oxides of Bismuth.

Like nitrogen, bismuth has four oxides:

Name.	M. F.	M. W.	G. F.
1. Bismuth dioxide	Bi_2O_2	450	
2. Bismuth oxide	Bi_2O_3	466	
3. Bismuth tetraoxide	Bi_4O_4	482	
4. Bismuth pentaoxide	Bi_2O_5	498	

The salts of Bi_2O_3 are the most important.

A concentrated solution of a *normal* salt of bismuth, poured into much water, partially decomposes and precipitates the *sub-* or *oxy*-salt.

The **structure** of these salts is not certainly known; they are *assumed* to be salts of the univalent radical *bismuthyl*, BiO. For this reason the molecular formula, molecular weight, and graphic formula of each of the *following bismuthyl* salts have been omitted.

Bismuthyl Nitrate (Subnitrate of Bismuth).

PREPARATION.—As a heavy, white, tasteless powder, very sparingly soluble in water, soluble in most acids, by dissolving metallic bismuth in nitric acid, and pouring this bismuth nitrate solution into much water:

$$Bi + 4HNO_3 = Bi(NO_3)_3 + NO + 2H_2O;$$

$$Bi(NO_3)_3 + 2H_2O = BiONO_3.H_2O + 2HNO_3.$$

The nitric acid of the latter decomposition keeps some bismuth in solution.

Bismuthyl Carbonate (Subcarbonate or Oxycarbonate of Bismuth, Pearl White).

PREPARATION.—This white or yellowish-white powder, precipitates with evolution of carbon dioxide when an alkali carbonate is added to bismuth nitrate solution:

$$2(Bi(NO_3)_3) + 3K_2CO_3 + H_2O = 6KNO_3 + 2CO_2 + (BiO)_2CO_3.H_2O.$$

On heating it loses water and carbon dioxide, forming the yellow oxide of bismuth, Bi_2O_3.

Bismuth Iodide (Subiodide of Bismuth).

PREPARATION.—Precipitates as a red powder by pouring a solution of bismuth nitrate in acetic acid into a solution of potassium iodide and sodium acetate: the latter prevents the formation of free nitric acid, which would decompose the subiodide:

$$2Bi(NO_3)_3 + 2H_2O + 2KI + 4NaC_2H_3O_2 = 2BiOI + 4NaNO_3 + 2KNO_3 + 4C_2H_3O_2 + 4H.$$

MERCURY (QUICKSILVER).

S., Hg; **V.**, 2; **A. W.**, 200.

OCCURRENCE.—The only liquid metal. In nature, as such, in small quantities, but usually as cinnabar (HgS).

PREPARATION.—From the sulphide, (1) by roasting or (2) distilling with lime.

PROPERTIES.—Freezes at —40° C.; boils at 357° C.; slightly volatile at all temperatures, silver white and lustrous; Sp. G. 13.56. The vapor is only one hundred times heavier than hydrogen, although the molecular weight is 200. Therefore the molecule of gaseous mercury is supposed to contain *one* atom.

It is bivalent, and forms *mercurous* and *mercuric* salts. The structure of mercurous and mercuric chloride is assumed to be

1. Mercuric chloride, $HgCl_2$, 2. Mercurous chloride, Hg_2Cl_2,

$$Hg{<}^{Cl}_{Cl}; \qquad \begin{array}{l} Hg-Cl \\ | \\ Hg-Cl \end{array}.$$

Mercury is oxidized in the air, insoluble in hydrochloric and dilute sulphuric acids, soluble in nitric and concentrated sulphuric acids, and combines directly with iodine, chlorine, and bromine.

USES.—As metal in making scientific instruments (thermometers, barometers, etc.), and amalgams (tin amalgams for looking-glasses); in preparing mercury and chalk, mercurial ointment, mercurial plaster, and blue mass (U. S. P.), and the various compounds of mercury.

COMPOUNDS OF MERCURY.

Mercurous Oxide (Black Oxide of Mercury).

M. F., Hg_2O; **M. W.**, 416; **G. F.**, $\begin{array}{l} Hg \\ | \\ Hg \end{array}{>}O$.

PROPERTIES AND PREPARATION.—As an insoluble black powder by the action of hydroxides of alkalies or alkaline earths on mercurous compounds:

$$Hg_2Cl_2 + Ca(OH)_2 = CaCl_2 + H_2O + Hg_2O.$$

Black Wash, U. S. P., is formed by mixing lime water and mercurous chloride.

Mercuric Oxide.

M. F., HgO; **M. W.**, 216; **G. F.**, $Hg{=}O$.

FORMS.—Two, differing in molecular structure according to their source:

1. **Yellow Mercuric Oxide,** U. S. P.—PREPARATION.—By mixing solutions of mercuric chloride and of hydroxide of an alkali or alkaline earth metal:

$$HgCl_2 + 2NaOH = 2NaCl + H_2O + HgO.$$

The orange-yellow, heavy precipitate is dried at 30° C.

2. **Red Mercuric Oxide,** U. S. P.—PREPARATION.—As a compact, crystalline substance, by heating the nitrate, or a mixture of the nitrate and metallic mercury:

$$Hg(NO_3)_2 + Hg = 2HgO + 2NO_2.$$

Yellow Wash, U. S. P., is a suspension of yellow oxide of mercury formed by adding mercuric chloride to lime water.

Mercurous Chloride (Calomel, Mild Chloride of Mercury, Subchloride or Protochloride of Mercury (U. S. P.)).

$$\textbf{M. F.},\ Hg_2Cl_2;\ \textbf{M. W.},\ 471;\ \textbf{G. F.},\ \begin{array}{l}Hg\text{—}Cl\\ |\\ Hg\text{—}Cl\end{array}.$$

PREPARATION.—By subliming a mixture of mercuric sulphate with a quantity of mercury equal to that contained in the sulphate and sodium chloride:

$$HgSO_4 + Hg + 2NaCl = Na_2SO_4 + Hg_2Cl_2;$$

and washing with *hot* water to remove remaining mercuric chloride.

PROPERTIES.—This heavy, white, insoluble, tasteless powder, volatilizes without fusing, easily decomposes by mineral acids, and alkali iodides, bromides, hydroxides, and carbonates, *which should not be administered with it unless the action of the decomposition-products is desired.*

Mercuric Chloride (Corrosive Chloride of Mercury, Corrosive Sublimate, Perchloride or Bichloride of Mercury, U. S. P.).

$$\textbf{M. F.},\ HgCl_2;\ \textbf{M. W.},\ 271;\ \textbf{G. F.},\ Hg{<}^{Cl}_{Cl}.$$

PREPARATION.—By subliming a mixture of mercuric sulphate and sodium chloride:

$$HgSO_4 + 2NaCl = HgCl_2 + Na_2SO_4.$$

PROPERTIES.—This white powder, soluble in water, alcohol, ether, and glycerin, forms heavy, colorless crystals, fuses at 265° C., volatilizes at 295° C., and has metallic taste, acid reaction, strong antiseptic and poisonous properties.

Mercurous Iodide (Yellow Iodide, Green Iodide, Protoiodide of Mercury, U. S. P.).

$$\textbf{M. F.},\ Hg_2I_2;\ \textbf{M. W.},\ 654;\ \textbf{G. F.},\ \begin{array}{l}Hg\text{—}I\\ |\\ Hg\text{—}I\end{array}.$$

PREPARATION.—1. By direct combination of the elements. 2. By precipitating mercurous solutions with potassium iodide.

PROPERTIES.—This yellow powder, insoluble in water, is unstable, easily decomposing into mercuric iodide and mercury. Exposure to light during its preparation soon causes its decomposition.

Mercuric Iodide (Red Iodide or Biniodide of Mercury, U. S. P.).

M. F., HgI_2; **M. W.**, 454; **G. F.**, $Hg<^{I}_{I}$.

PREPARATION.—As a red powder, by adding potassium iodide to mercuric chloride. It dissolves in excess of either salt, forming a colorless solution.

Mercurous Sulphate.—

$$\textbf{M. F.},\ Hg_2SO_4;\ \textbf{M. W.},\ 496;\ \textbf{G. F.},\ SO_2\left\langle\begin{array}{l}O-Hg\\ \quad\ \ |\\ O-Hg\end{array}\right..$$

PREPARATION.—By triturating mercuric sulphate with mercury.

Mercuric Sulphate.

M. F., $HgSO_4$; **M. W.**, 296; **G. F.**, $SO_2<^{O}_{O}>Hg$.

PREPARATION.—As a heavy white crystalline powder by dissolving mercury in concentrated sulphuric acid:

$$Hg + 2H_2SO_4 = HgSO_4 + 2H_2O + SO_2.$$

Yellow Mercuric Subsulphate (Basic Mercuric Sulphate, Turpeth Mineral, Mercuric Oxysulphate, U. S. P.).—

M. F., $HgSO_4(HgO)_2$; **M. W.**, 727.4; **G. F.**, not determined.

PREPARATION.—As a lemon-yellow tasteless powder, almost insoluble in water, by adding mercuric sulphate to boiling water.

Mercurous Nitrate.—

$$\textbf{M. F.},\ Hg_2(NO_3)_2;\ \textbf{M. W.},\ 524;\ \textbf{G. F.},\ \begin{array}{l}Hg-ONO_2\\ |\\ Hg-ONO_2\end{array}.$$

FORMED by dissolving mercury in nitric acid, with excess of mercury.

Mercuric Nitrate.—

M. F., $Hg(NO_3)_2$; **M. W.**, 324; **G. F.**, $Hg<^{ONO_2}_{ONO_2}$.

FORMED when excess of nitric acid acts on mercury. Both nitrates are decomposed by water, forming basic salts.

Mercuric Sulphide.

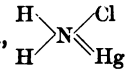
M. F., HgS; **M. W.**, 232; **G. F.**, $Hg{=}S$.

OCCURRENCE.—In nature as cinnabar, a red mineral.

PREPARATION.—When hydrogen sulphide acts on mercuric salts it forms *black* mercuric sulphide. By sublimation of this the crystalline *red* sulphide (*vermilion* or *cinnabar*) is made.

Mercurous and **mercuric sulphides** form also by direct union of the elements in proper proportions.

Ammoniated Mercury (White Precipitate, U. S. P.).

M. F., NH_2HgCl; **M. W.**, 251.2; **G. F.**,  .

OBTAINED as a white powder by pouring a solution of mercuric chloride into ammonia water:

$$HgCl_2 + 2NH_4OH = NH_2HgCl + NH_4Cl + 2H_2O.$$

PROPERTIES.—This metallammonium compound may be looked on as ammonium chloride in which *two* atoms of hydrogen have been replaced by mercury.

The Toxicology of Mercury.

Very many compounds of mercury are poisonous.

CHEMICAL ANTIDOTE.—White of egg (albumin) should be given (in small quantities at a time to avoid redissolving the precipitate first formed) and promptly followed by an emetic or lavage.

SILVER.

S., Ag; **V.**, 1; **A. W.**, 108.

OCCURS in nature occasionally as the metal, more often as sulphide, usually combined with lead sulphide (*argentiferous galena*).

PREPARATION.—By roasting this ore in a current of air. The lead is converted to litharge and the silver is freed.

PROPERTIES.—Silver is the whitest metal, takes the highest polish, is the best conductor of heat and electricity, is not oxidized in the air, but is coated with a film of black sulphide in air containing hydrogen sulphide, dissolves in nitric and sulphuric acids, but not in hydrochloric acid, is univalent, and forms *one* series of salts.

ALLOTROPIC MODIFICATIONS OF SILVER.

Silver is the only metal with allotropic modifications, which are chiefly obtained by reduction of its salts, and differ from ordinary silver—(1) by having various colors (blue, greenish blue, red, purple, yellow); (2) by being converted into chloride by hydrochloric acid; (3) by solubility

in ammonia; and (4) by reducing certain oxidizing agents. They are converted into ordinary silver by strong acids, heat, electricity, and other forms of energy.

Alloys of silver and copper, used for coin, contain 5–25 per cent. of copper, for hardness.

COMPOUNDS OF SILVER.

Silver Nitrate, U. S. P.

M. F., $AgNO_3$; **M. W.**, 170; **G. F.**, $\begin{matrix} O \\ O \end{matrix}\!\!>\!N{-}O{-}Ag$.

PROPERTIES.—Forms colorless, transparent crystals, soluble in less than their own weight of water. Exposed to light, or in contact with animal matter, silver nitrate decomposes and blackens, because of the separation of metallic silver. The caustic action of silver nitrate is due to its separation into silver and nitric acid by the tissues.

PREPARATION.—By dissolving silver in nitric acid and crystallizing the solution. When prepared from coin alloy, the cupric nitrate formed is removed by evaporating the solution to dryness and fusing the residue; thus the cupric nitrate is converted into cupric oxide, and the silver salt is dissolved from the fused mass and the solution crystallized.

USES.—As a caustic, and largely in photography, and in the manufacture of indelible ink and hair-dye.

Moulded Silver Nitrate (Caustic, Lunar Caustic, Lapis Infernalis, U. S. P.).—OBTAINED by fusing a mixture of silver nitrate and 4 per cent. hydrochloric acid, and pouring it into moulds to form white cylindrical sticks.

Diluted Silver Nitrate (Mitigated Caustic, U. S. P.).—OBTAINED by fusing silver nitrate with twice its weight of potassium nitrate.

Ammonio-Silver Nitrate is formed by adding enough ammonia to silver nitrate solution to dissolve the precipitate first formed.

Silver Oxide.

M. F., Ag_2O; **M. W.**, 232; **G. F.**, $\begin{matrix} Ag \\ Ag \end{matrix}\!\!>\!O$.

PROPERTIES.—A dark-brown powder, sparingly soluble in water, with weak alkaline reaction, is a strong base, easily decomposes into silver and oxygen.

PREPARATION.—By the action of an alkali hydroxide on silver nitrate:

$$2AgNO_3 + 2NaOH = 2NaNO_3 + H_2O + Ag_2O.$$

Silver Iodide.

M. F., AgI; **M. W.**, 235; **G. F.**, Ag—I.

PROPERTIES.—An amorphous yellowish powder, insoluble in water, slightly so in ammonium hydroxide.

PREPARATION.—By decomposition of silver nitrate by potassium iodide:

$$AgNO_3 + KI = KNO_3 + AgI.$$

Silver Bromide.

M. F., $AgBr$; **M. W.**, 188; **G. F.**, Ag—Br.

Silver Chloride.

M. F., $AgCl$; **M. W.**, 143.5; **G. F.**, Ag—Cl.

PREPARATION.—By analogous decomposition. Silver chloride is insoluble in water and acids, soluble in ammonia water. Actinic rays of light decompose halogen compounds of silver, changing their color to violet, brown, and finally black, the iodide and bromide being more sensitive than the chloride.

CADMIUM AND COMPOUNDS OF CADMIUM.

S., Cd; **V.**, 2; **A. W.**, 112.

PROPERTIES.—Resembles zinc *physically* and usually occurs in nature associated with zinc.

Cadmium Sulphide.—

M. F., CdS; **M. W.**, 144; **G. F.**, Cd=S.

A yellow compound used as a pigment.

Cadmium Iodide.—

M. F., CdI_2; **M. W.**, 366; **G. F.**, $Cd<^{I}_{I}$.

Like the sulphide is sometimes used in medicine.

Qualitative Analysis of the Lead Group.

These metals may be divided into *two* classes, by the behavior of their compounds on addition of ammonium hydroxide. Lead, bismuth, and mercury form precipitates insoluble in excess of the reagent, but the copper, silver, and cadmium precipitates are soluble. The analytical

reactions of each of these two classes and of mercuric and mercurous compounds are summarized in the following tables.

Subclass I.

Reagents.	Group Members.			
	Lead.	Bismuth.	Mercuric Salts.	Mercurous Salts.
H_2S	Black precipitate	Dark-brown precipitate	Black precipitate	Black precipitate
KOH	White precipitate	White precipitate	Yellow precipitate	Black precipitate
NH_4OH	White precipitate	White precipitate	White precipitate	Black precipitate
In excess of reagent	Insoluble	Insoluble	Insoluble	Insoluble
Na_2CO_3	White precipitate	White precipitate	Reddish-brown precipitate	Yellowish precipitate
KI	Yellow precipitate	Brown precipitate	Red precipitate	Yellowish-green precipitate
In excess of reagent	Sparingly soluble	Insoluble	Soluble	Partly soluble
K_2CrO_4	Yellow precipitate	. . . .	Orange precipitate	Red precipitate
HCl	White precipitate soluble in hot water	. . . .	. . . .	White precipitate darkens with ammonia
H_2SO_4	White precipitate	. . . .	. . . .	White precipitate

Subclass II.

Reagents.	Group Members.		
	Copper.	Silver.	Cadmium.
H_2S	Black precipitate	Black precipitate	Yellow precipitate
KOH	Blue precipitate brown on boiling	Brown precipitate	White precipitate
NH_4OH	Pale-blue precipitate	Brown precipitate	White precipitate
In excess of reagent	Dark-blue solution	Colorless solution	Colorless solution
Na_2CO_3	Greenish-blue precipitate	Pale-yellow precipitate	White precipitate
KI	Yellow precipitate	Pale-yellow precipitate	
K_2CrO_4	Orange precipitate	Dark-red precipitate	
HCl or soluble chlorides	. . . .	White precipitate soluble in ammonia	

QUESTIONS.

How do the metals of the lead group differ from the iron and arsenic groups? How do they act on sulphuric and nitric acids?

State symbol, valence, physical and chemical properties of lead.

How may a lead pipe contaminate drinking-water?

How may the presence of lead in water be demonstrated?

State composition, properties, and mode of manufacture of litharge and red lead.

In what property is lead nitrate different from the other lead salts of mineral acids?

How is lead carbonate formed, and what are its properties?

How is lead iodide made, what are its properties, and how is it distinguished from lead chromate?

State occurrence in nature, and physical and chemical properties of copper?

How many series of salts does copper form?

State composition, mode of manufacture, properties, and uses of black oxide of copper and cuprous oxide.

How is blue vitriol obtained? State its composition and properties.

How is anhydrous cupric sulphate obtained?

What are metallammoniums?

Mention name and composition of the ammonio-cupric sulphates.

What is the valence of bismuth, and what element does it resemble in its compounds with oxygen?

What is meant by bismuthyl compounds, and how are they obtained?

Show by equation the formation of bismuthyl nitrate.

How does mercury occur in nature?

State physical and chemical properties and uses of mercury.

What is the valence of mercury, and how many series of salts does it form?

State composition, mode of manufacture, and properties of the three oxides of mercury.

What are black wash and yellow wash?

How is mercuric chloride made, and what are its properties?

How is mercurous chloride made?

State composition, mode of manufacture, and properties of mercurous and mercuric iodides.

How is mercuric sulphate made, and what are its properties?

What is turpeth mineral, and how is it made?

State composition and mode of manufacture of the nitrates of mercury.

What is vermilion?

What are the composition and mode of manufacture of white precipitate?

State physical and chemical properties of silver.

What is coin alloy?

How is silver nitrate made from pure silver, and from coin?

What are lunar caustic, and mitigated caustic?

How may mercurous be distinguished from mercuric salts?

State reactions for distinguishing the metals of the lead group from each other.

CHAPTER XX.

THE ARSENIC GROUP.

ARSENIC, Antimony, Tin, Gold, Platinum, and Molybdenum comprise this group.

GROUP PROPERTIES.—Hydrogen sulphide added to a solution of any of them precipitates the sulphide, insoluble in dilute acids, soluble in ammonium sulphide or alkali hydroxides, due to the formation of soluble double compounds.

The properties of the sulphides and the tendency of the oxides to form acids are the chief points of similarity between these metals. Arsenic and antimony are more closely related than any of the others. They resemble the non-metals nitrogen and phosphorus in valence, being either tri- or pentavalent, and in several compounds, arsenic resembles phosphorus in containing four atoms to the molecule.

The sulphides of arsenic are yellow, of antimony orange, of tin brown or yellow, of gold and platinum brown in color.

ARSENIC.

S., As ; **V.**, 3 or 5 ; **A. W.**, 75.

OCCURRENCE.—In nature chiefly as mispickel (FeSAs), realgar (As_2S_2), and orpiment or auripigment (As_2S_3).

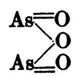
PROPERTIES.—It is steel-gray, brittle, odorless, volatilizes without fusing at 180° C. if air is excluded; burns with a bluish-white light in the air, forming arsenous oxide, is easily converted into the oxide by exposure to air (for this reason water digested with arsenic will soon contain arsenous oxide, formed by the action of the oxygen absorbed in the water).

USES.—As metal in alloys and fly-poison.

COMPOUNDS OF ARSENIC.

Arsenous Oxide (White Arsenic, Arsenic Trioxide, Arsenous Anhydride, Rats' Bane).

M. F., As_2O_3; **M. W.**, 198; **G. F.**,

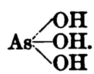


PREPARATION.—As a by-product in the roasting of ores of other metals containing arsenic, when it volatilizes, and is collected in chambers or long flues.

PROPERTIES.—A white, crystalline powder, or a transparent mass, volatile at 218° C. without previously melting; recently sublimed (*vitreous arsenous oxide*), it is a transparent mass, gradually becoming opaque, by a rearrangement of the molecules into crystals, which may be seen under the microscope.

Arsenous Acid.

M. F., H_3AsO_3; **M. W.**, 136; **G. F.**,



PREPARATION.—By dissolving arsenous oxide in water:

$$As_2O_3 + 3H_2O = 2H_3AsO_3.$$

The acid is known only in solutions in water.

Metarsenous Acid.—

M. F., $HAsO_2$; **M. W.**, 108; **G. F.**, $As\begin{array}{l}\diagup\!\!\diagup O\\ \diagdown OH\end{array}$

OCCURRENCE.—Known only in salts—*e. g.*, sodium metarsenite.
PREPARATION.—By the action of arsenous oxide on sodium hydroxide:

$$2NaOH + As_2O_3 = 2NaAsO_2 + H_2O.$$

Liquor Acidi Arsenosi, U. S. P., is a 1 per cent. solution of As_2O_3 in water containing 0.5 per cent. of dilute hydrochloric acid.

Liquor Potassii Arsenitis (Fowler's Solution), U. S. P., is a solution of potassium metarsenite made by dissolving 1 per cent. arsenous acid and 2 per cent. potassium bicarbonate in water containing 3 per cent. compound tincture of lavender.

Arsenic Oxide (Arsenic Pentoxide, Anhydrous Arsenic Acid).

M. F., As_2O_5; **M. W.**, 230; **G. F.**, $\begin{array}{l}\quad\diagup\!\!\diagup O\\ As{=}O\\ \quad\diagdown\\ \qquad O.\\ \quad\diagup\\ As{=}O\\ \quad\diagdown\!\!\diagdown O\end{array}$

PREPARATION.—By expelling water by heat from arsenic acid:

$$2H_3AsO_4 = As_2O_5 + 3H_2O.$$

PROPERTIES.—Resembles phosphoric oxide closely, uniting with water in varying proportions into ortho-, pyro-, and metarsenic acid. The last two, dissolved in water, pass at once into:

Orthoarsenic Acid (Arsenic Acid).—

M. F., H_3AsO_4; **M. W.**, 142; **G. F.**, $O{=}As\begin{array}{l}\diagup O{-}H\\ -O{-}H.\\ \diagdown O{-}H\end{array}$

PREPARATION.—By heating arsenous oxide with nitric acid:

$$3As_2O_3 + 4HNO_3 + 7H_2O = 6H_3AsO_4 + 4NO.$$

Disodium-Hydrogen Arsenate (Sodium Arsenate).

M. F., $Na_2HAsO_4.7H_2O$; **M. W.**, 311.9; **G. F.**, $\begin{array}{r}NaO\diagdown\\ NaO\,\rangle\\ OH\diagup\end{array}As{=}O.$

PREPARATION.—By fusing arsenous oxide with sodium carbonate

and nitrate, and dissolving the sodium pyroarsenate so formed in water and crystallizing:

$$As_2O_3 + 2NaNO_3 + Na_2CO_3 = Na_4As_2O_7 + N_2O_3 + CO_2;$$

$$Na_4As_2O_7 + 15H_2O = 2(Na_2HAsO_4.7H_2O).$$

Liquor Sodii Arsenatis, U. S. P., is a 1 per cent. solution in water.

Hydrogen Arsenide (Arsine, Arsenuretted Hydrogen).

M. F., AsH_3; **M. W.**, 78; **G. F.**, $As\left\langle\begin{matrix}H\\H\\H\end{matrix}\right.$.

PREPARATION.—By nascent hydrogen acting on the oxides of arsenic, arsenic acid, or their salts.

PROPERTIES.—A colorless, very poisonous gas, with an odor like garlic, and burning to white clouds of arsenous oxide with a bluish flame whose structure resembles the ordinary flame, the *carbon being replaced by arsenic: therefore, a deposit of metallic arsenic occurs on a cold plate of porcelain held in this flame.*

When passed through a red-hot tube arsenic oxide is decomposed, the arsenic being deposited as a bright metallic ring in the cooler part of the tube.

Arsenic Disulphide (Realgar); Arsenic Trisulphide (Orpiment).

M. F., As_2S_2, As_2S_3; **M. W.**, 213.8, 245.8;

G. F., $\begin{matrix}As=S\\ |\\ As=S\end{matrix}$; $\begin{matrix}As=S\\ \\ As=S\end{matrix}\!\!>\!S$, respectively.

PROPERTIES.—The disulphide is orange-yellow, and the trisulphide is golden-yellow, and both are fusible and volatile substances, and combine with other metallic sulphides into double salts called *sulpharsenides.*

Arsenous Iodide.

M. F., AsI_3; **M. W.**, 456; **G. F.**, $As\left\langle\begin{matrix}I\\I\\I\end{matrix}\right.$.

PREPARATION.—By direct combination of the elements.

Donovan's Solution, U. S. P., is a watery solution of 1 per cent. *each of arsenous* iodide and mercurous iodide.

Toxicology of Arsenic.

Arsenic and its compounds are poisonous.

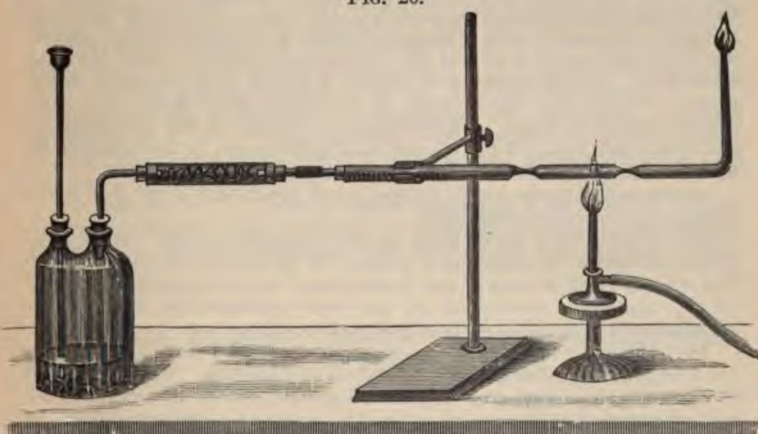
CHEMICAL ANTIDOTES.—Moist, freshly prepared ferric hydroxide, or ferric hydrate and magnesia, or dialyzed iron, forms insoluble ferric arsenite or arsenate. Prompt emesis or lavage should follow.

TESTS FOR ARSENIC.—In cases of suspected poisoning by arsenic, the stomach contents, or other organic material (liver, kidneys, etc.), should be cut into pieces, well mashed and mixed with water, and heated over a water-bath with potassium chlorate and hydrochloric acid until the mixture is uniformly yellow, and all free chlorine eliminated. The clear solution is treated with hydrogen sulphide gas for several hours, the precipitate is heated with ammonium sulphide solution, and this solution filtered, and the filtrate evaporated to dryness. The residue so obtained is fused with sodium carbonate and nitrate, and the fused mass dissolved in warm water. With this final solution the following tests are made:

I. *Marsh's test* permits the collection of the total arsenic present, and is therefore a *quantitative* and *qualitative* test.

(*a*) *Qualitative Marsh's Test.*—The *Marsh apparatus* (Fig. 20) con-

FIG. 20.



Marsh's apparatus for detection of arsenic.

sists of a glass vessel, provided with a funnel-tube and a right-angle delivery-tube connecting it with a wider tube containing calcium chloride or some other drying material. This tube is connected with a combustion-tube about 40 cm. long, about 1 cm. in diameter, and drawn out at intervals of 10 cm. to half this diameter.

Hydrogen is generated in the glass vessel by the action of dilute sulphuric acid on zinc, and, after the air has been driven from the apparatus, is ignited at the end of the tube. The suspected solution is now passed into the vessel through the funnel-tube and the flame assumes the characteristics of burning arsenic. *The spot deposited on a plate must be differentiated from similar spots deposited from burning antimoniated hydrogen:* (1) by treating the spots with solution of a hypochlorite. When *arsenic* the spots *dissolve*, but *antimony* spots are *unchanged*; (2) by dissolving the spots with nitric acid, evaporating to dryness, and moistening the residue with silver nitrate, *arsenic* spots *turn red*, *antimony* spots *remain white.*

(*b*) *Quantitative Marsh's Test.*—The glass tube is heated to redness at two of its wide parts so as to decompose the arsenuretted hydrogen. The arsenic solution should be added slowly so as to avoid the production of more arsenuretted hydrogen at a time than can be decomposed. To avoid loss of arsenic by overproduction, the end of the tube is inverted into a solution of silver nitrate, which reacts with arsenic, forming silver and arsenous acid.

II. *Gutzeit's test* is made in a test-tube provided with a cap of three thicknesses of filter-paper. Hydrogen is generated by zinc and dilute sulphuric acid, and a few drops of the suspended solution added and the cap fastened over the mouth of the tube. The upper paper of the cap is moistened with a saturated solution of silver nitrate containing 1 per cent. nitric acid and the apparatus placed in a dark place for some time. In the presence of arsenic a yellow stain develops on the paper. When the yellow stain is moistened with water the color becomes brown or black. The yellow color is due to the formation of the compound $Ag_3As(AgNO_3)_3$, which is decomposed by the addition of water, forming metallic silver, arsenous and nitric acids.

Antimonetted hydrogen in Gutzeit's test produces a dark spot on the paper, but causes no yellow color.

III. *Fleitmann's test* is made as the foregoing, except that the hydrogen is evolved by action of a solution of an alkali hydroxide on zinc. The stain is black or brown, and is developed at a temperature of about 90° C. *Antimony compounds do not form antimonetted hydrogen in alkaline solutions and consequently do not interfere.*

IV. *Reinsch's test* is made by immersing a bright piece of copper in the acidified, suspected solution of arsenic and heating. The copper will be coated with a steel-gray deposit of arsenic, which is volatilized by heat, forming a sublimate of octahedral crystals. *Antimony forms a bluish deposit on copper, and when volatilized forms an amorphous sublimate.*

ANTIMONY (STIBIUM).

S., Sb; **V.**, 3 or 5; **A. W.**, 120.

OCCURRENCE.—In nature as the trisulphide, Sb_2S_3, stibnite, **black** *antimony, or crude* antimony.

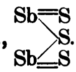
Preparation.—By roasting the ore, and reducing with charcoal.

Properties.—This bluish-white metal fuses at 450° C., volatilizes at higher temperatures without the access of air. In air it burns when heated.

Uses.—In many alloys.

COMPOUNDS OF ANTIMONY.

Antimonous Sulphide (Antimony Trisulphide).

M. F., Sb_2S_3; **M. W.**, 336; **G. F.**, 

Occurrence.—In nature usually mixed with other ores.

Preparation.—(1) From these ores by heating. The antimonous sulphide melts and is poured off, cooling as steel-gray masses, or a grayish-black lustreless powder, or as an orange powder. (2) By adding sulphuretted hydrogen to an antimonous solution. On heating, this orange sulphide becomes black sulphide.

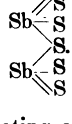
Glass of antimony or *crocus* is obtained by roasting the native sulphide in air until partially decomposed and fused into a vitreous mass.

Antimonii Sulphidum Purificatum (U. S. P.).—**Preparation.**—By treating powdered antimonous sulphide with ammonia water for five days, washing and drying, thus removing all traces of arsenic.

Sulphurated Antimony (Oxysulphide of Antimony, Kermes Mineral, U. S. P.).—**Preparation.**—By dissolving purified antimonous sulphide in sodium hydroxide solution and adding sulphuric acid as long as a precipitate forms.

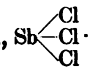
Properties.—This reddish-brown powder is a mixture of antimonous sulphide and a little antimonous oxide, in inconstant proportions; therefore, no molecular formula, molecular weight, and graphic formula may be given.

Antimonic Sulphide (Golden Sulphuret of Antimony, Antimony Pentasulphide).

M. F., Sb_2S_5; **M. W.**, 400; **G. F.**, 

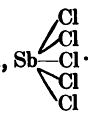
Preparation.—As a red powder, by precipitating acid solution of antimonic salts with sulphuretted hydrogen. Antimonous and antimonic sulphides and oxides, like similar arsenic compounds, unite with other metallic salts to form sulpho- or oxy-salts.

Antimonous Chloride (Butter of Antimony).

M. F., $SbCl_3$; **M. W.**, 226.5; **G. F.**, 

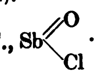
PREPARATION.—As a white crystalline, semitransparent mass, melting at 74° C. into a yellow oily liquid, by dissolving native sulphide in hydrochloric acid, evaporating the solution, and distilling the chloride.

Antimonic Chloride.

M. F., $SbCl_5$; **M. W.**, 297.5; **G. F.**, 

PREPARATION.—As a fuming liquid by passing chlorine over antimonous chloride.

Antimonyl Chloride (Powder of Algaroth).—

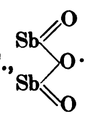
M. F., SbOCl; **M. W.**, 171.5; **G. F.**, 

PREPARATION.—With some antimonous oxide, by pouring a concentrated solution of antimonous chloride in much water:

$$3SbCl_3 + 4H_2O = SbOCl + Sb_2O_3 + HCl.$$

STRUCTURE.—Analogous to bismuthyl salts, being salts of the *radical antimonyl* (SbO).

Antimonous Oxide.

M. F., Sb_2O_3; **M. W.**, 288; **G. F.**, 

PREPARATION.—As a heavy grayish-white powder, by treating the foregoing precipitate with sodium carbonate:

$$2SbOCl + Na_2CO_3 = Sb_2O_3 + 2NaCl + CO_2.$$

PROPERTIES.—Insoluble in water, soluble in hydrochloric acid and warm solution of tartaric acid.

Tartar Emetic (Potassium Antimonyl Tartrate, U. S. P.).— PREPARATION.—By dissolving the moist and freshly precipitated oxide *in potassium* bitartrate solution.

Toxicology of Antimony.

The salts are poisonous.

CHEMICAL ANTIDOTES.—Tannic acid or freshly prepared ferric hydroxide. The salts of antimony usually cause vomiting, but if this does not occur, lavage must be done.

TESTS.—See Arsenic.

TIN (STANNUM).

S., Sn; **V.**, 2; **A. W.**, 119.

OCCURRENCE.—In nature chiefly as cassiterite or tin-stone, SnO_2.

PREPARATION.—By heating this ore with carbon. The *oxide* (*putty powder*) is used for polishing.

PROPERTIES.—A silver-white metal, fusing at 228° C., and used in many alloys. *Tin-plate* is sheet iron covered with a thin coating of tin. *Tin-amalgam* is used for silvering looking-glasses.

COMPOUNDS OF TIN.

There are two series of compounds, *stannous* having the metal bivalent, and *stannic* having it quadrivalent.

Stannous Chloride (Protochloride of Tin).

M. F., $SnCl_2$; **M. W.**, 225.5; **G. F.**, $Sn{<}^{Cl}_{Cl}$.

PREPARATION.—By dissolving tin in hot hydrochloric acid. Is a strong reducing agent, and used as a reagent and mordant in dyeing. Poured into much water it forms a precipitate of *oxychloride of tin.*

Stannic Chloride (Perchloride of Tin).

M. F., $SnCl_4$; **M. W.**, 261; **G. F.**, $Sn\begin{cases}Cl\\Cl\\Cl\\Cl\end{cases}$.

PREPARATION.—1. By passing chlorine through a solution of stannous chloride. 2. By treating it with aqua regia.

Stannic Sulphide.

M. F., SnS_2; **M. W.**, 183; **G. F.**, $Sn{\lll}^{S}_{S}$.

PREPARATION.—Precipitated from stannic solutions by hydrogen sul-

phide. *Mosaic gold*, or *bronze powder*, is stannic sulphide, a golden-colored powder, obtained by the action of sulphur on tin amalgam and ammonium chloride.

GOLD (AURUM).

S., Au; **V.**, 4; **A. W.**, 197.

OCCURRENCE.—In nature chiefly in the free state, often mixed with other metals.

PREPARATION.—1. By washing ores rich in gold, separating by mechanical means and then dissolving in mercury. The amalgam is decomposed by heating in a retort, the mercury distilling and gold remaining. 2. From ores containing little gold, by treating the finely powdered material with potassium cyanide solution. The gold dissolves as double cyanide of gold and potassium, $AuK(CN)_2$, and precipitates from the solution by electrolysis.

Chemically pure gold is obtained by dissolving gold in aqua regia, removing free acid from the solution by evaporation, precipitating platinum by adding alcohol and potassium chloride, and from the filtrate precipitating gold by adding a reducing agent, as oxalic acid, ferrous sulphate, etc. The precipitate is without lustre and is converted into a metallic mass by fusion.

PROPERTIES.—Gold is orange-yellow, green by transmitted light, has great ductility and malleability. In forming foil its cohesiveness is greatly diminished. By heating the foil to redness its cohesiveness is restored. Gold is soluble in aqua regia, free chlorine, and mercury. It is trivalent in *auric* compounds, and univalent in *aurous*.

COMMERCIAL FORMS.—*Pure gold* is too soft for general use, and is therefore alloyed with silver and copper. The fineness of gold is expressed in *carats*, a term meaning one twenty-fourth part. Eighteen carats fineness indicates an alloy containing $\frac{18}{24}$ pure gold. American coin contains 90 per cent. of gold.

Refined gold is gold treated with nitric or sulphuric acid to remove metals like silver and copper.

COMPOUNDS OF GOLD.

Auric Chloride (Chloride of Gold).—

M. F., $AuCl_3$; **M. W.**, 303.5; **G. F.**, $Au{<}^{Cl}_{Cl}{-}Cl$.

PREPARATION.—By dissolving gold in aqua regia and evaporating.

Gold and Sodium Chloride is a mixture of equal parts of chloride of gold and chloride of sodium.

PLATINUM.

S., Pt; **V.**, 4; **A. W.**, 194.

OCCURRENCE.—In nature free, but usually associated with certain rare metals (*iridium*, *osmium*, *rhodium*, etc.) which resemble it.

PROPERTIES.—Being particularly resistant to heat and reagents, it is used for making chemical apparatus; is quadrivalent, and dissolves in aqua regia.

COMPOUNDS OF PLATINUM.

Platinic Chloride.—

M. F., $PtCl_4$; **M. W.**, 336; **G. F.**, $Pt\begin{cases}Cl\\Cl\\Cl\\Cl\end{cases}$.

PREPARATION.—By dissolving platinum in aqua regia.

USES.—As a reagent for potassium and ammonium compounds, forming insoluble double salts, $PtCl_4(KCl)_2$ or $PtCl_4(NH_4Cl)_2$.

MOLYBDENUM.

S., Mo; **V.**, 4; **A. W.**, 96.

OCCURRENCE.—In nature as molybdenite (MoS_2).

PREPARATION.—By roasting this ore molybdic oxide, MoO_3 forms, which unites with water into an acid whose salts are *molybdates*.

COMPOUNDS OF MOLYBDENUM.

Ammonium Molybdate.—

M. F., $(NH_4)_2MoO_4$; **M. W.**, 196; **G. F.**, $MoO<\begin{matrix}O-NH_4\\O-NH_4\end{matrix}$.

Dissolved in dilute nitric acid it is a very sensitive reagent for orthophosphoric acid and its salts (see page 122).

Phosphomolybdic Acid.—

M. F., $(H_3PO_4)_2 22MoO$; **M. W.**, 3362.

It is a *mixture*, as shown by the molecular formula, therefore no satisfactory graphic formula may be indicated.

USE.—A reagent, is prepared by digesting phosphoric acid with molybdic oxide until a clear solution is obtained, which is evaporated and crystallized.

Qualitative Analysis of the Arsenic Group.

The chief analytical reactions of arsenic and antimony have been indicated in the consideration of the detection of arsenic in organic matter. Tin is distinguished from gold and platinum by the solubility of its sulphide in strong hydrochloric acid. Platinum forms insoluble double salts with ammonium and potassium compounds in the presence of hydrochloric acid. Gold is easily precipitated from its solution by reducing agents, the dark precipitate being converted into the metallic form by heating.

QUESTIONS.

State the characteristics of the arsenic group of metals.
What are the colors of the sulphides of these metals?
Which non-metallic elements do arsenic and antimony resemble, and how?
How is metallic arsenic obtained from realgar?
What is the effect of water on metallic arsenic?
State the composition, mode of manufacture, and properties of arsenous oxide?
What is the composition of arsenous acid and of metarsenous acid?
What is the composition of Fowler's solution and of liquor acidi arsenosi?
State the composition and mode of manufacture of arsenic oxide and sodium arsenate.
State the composition, mode of manufacture, and properties of arsenic.
What are sulpharsenites?
What is Donovan's solution?
Describe Marsh's, Gutzeit's, Fleitmann's, and Reinsch's tests for arsenic.
How may antimony be distinguished from arsenic?
State the composition, mode of manufacture, and properties of the sulphides of antimony.
What is butter of antimony?
What is powder of Algaroth, and how is it obtained?
State the composition and properties of antimonous oxide.
State the occurrence in nature and uses of tin.
State the compositions of the chlorides of tin.
What is Mosaic gold?
How is gold separated from its ores?
How are chemically pure gold, refined gold, and cohesive gold obtained?
What is indicated by eighteen-carat gold?
What are the properties of platinum?
What is ammonium molybdate, and for what is it used?
How is phosphomolybic acid made?

APPENDIX.

IMPORTANT PHYSICAL AND CHEMICAL DATA.

TABLES OF WEIGHTS AND MEASURES.

Measures of Length.

Metric System.		U. S. Linear Measure.
1 millimeter = 0.001 meter	=	0.0394 inch.
1 centimeter = 0.01 meter	=	0.3937 inch.
1 decimeter = 0.1 meter	=	3.9371 inches.
1 meter	=	39.3708 inches.
1 decameter = 10 meters	=	32.8089 feet.
1 hectometer = 100 meters	=	328.089 feet.
1 kilometer = 1000 meters	=	0.6214 mile.
1 yard or 36 inches	=	0.9144 meter.
1 inch	=	25.4 millimeters.

I. Measures of Capacity.

Metric System.		U. S. Liquid Measure.
1 milliliter = 1 c.c. = 0.001 liter	=	0.0021 pint.
1 centiliter = 10 c.c. = 0.01 liter	=	0.0211 pint.
1 deciliter = 100 c.c. = 0.1 liter	=	0.2113 pint.
1 liter = 1000 c.c.	=	1.0567 quart.
1 decaliter = 10 liters	=	2.6418 gallons.
1 hectoliter = 100 liters	=	26.418 gallons.
1 kiloliter = 1000 liters	=	264.18 gallons.

II. Measures of Capacity.

U. S. Liquid Measure.		Metric System.
1 U. S. gallon	=	3785.3 c.c.
1 imperial gallon	=	4543.5 c.c.
1 minim	=	0.06 c.c.
1 fluidrachm	=	3.70 c.c.
1 fluidounce	=	29.57 c.c.
1 liter	=	33.81 fluidounces.

I. Measures of Weight.

Metric System.				Troy Weight.
1 milligram	=	0.001 gram	=	0.015 grain.
1 centigram	=	0.01 gram	=	0.154 grain.
1 decigram	=	0.1 gram	=	1.543 grain.
1 gram			=	15.432 grains.
1 decagram	=	10 grams	=	154.324 grains.
1 hectogram	=	100 grams	=	0.268 pound.
1 kilogram	=	1000 grams	=	2.679 pounds.

II. Measures of Weight.

Troy Weight.		Metric System.
1 grain	=	0.0648 gram.
1 drachm	=	3.888 grams.
1 ounce	=	31.103 grams.
1 ounce **avoirdupois**	=	28.350 grams.
1 pound **avoirdupois**	=	453.592 grams.
1 crith	=	0.0896 gram.

Commercial Weights and Measures of the U. S. A.

1 pound avoirdupois	=	16 ounces.
1 ounce	=	437.5 grains.
1 gallon	=	231 cubic inches.
1 gallon	=	4 quarts = 8 pints.

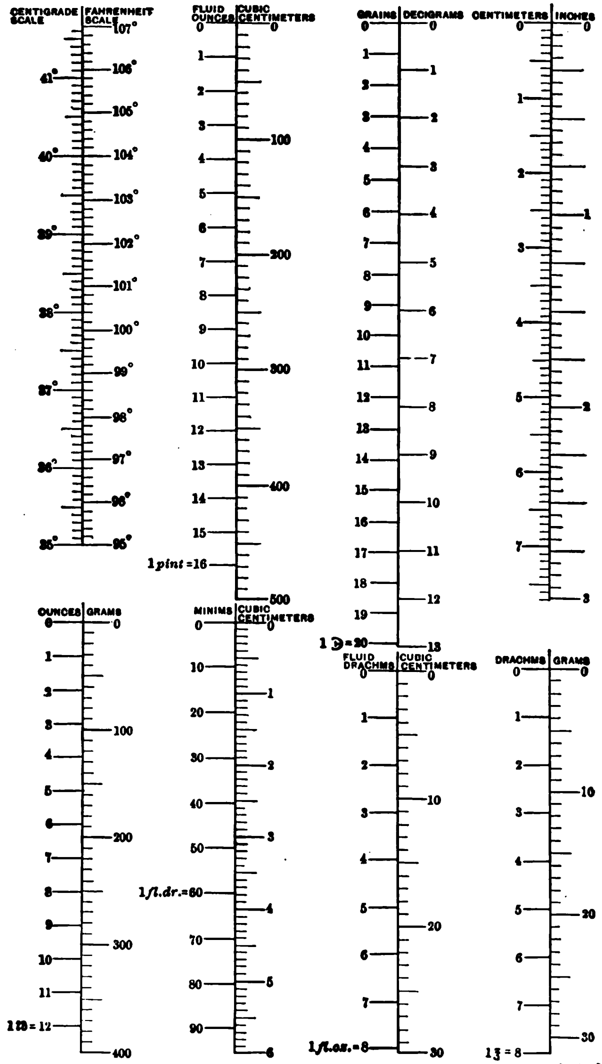
1 pint of water weighs 7291.2 grains at a temperature of 15.6° C.

Troy Weight.

1 drachm	=	60 grains.		
1 ounce	=	8 drachms	=	480 grains.

Diagram of Metric and Apothecary System Equivalents.

COMPARATIVE SCALES, showing at a glance the exact equivalent of ordinary weights and measures in those of the Metric System, and *vice versa.*



The equivalents of fractions, whether large or small, may be found with great nicety by *these scales. For instance,* 1/6 grain = 1/6 of the metric equivalent of 7 grains, and 1-200 grain =1-4000 *of the* metric equivalent of 20 grains. This method is, of course, reversible.

MEASURES OF ENERGY.

DYNE is the unit of force in the centimeter-gram-second system of measurement. It is that force which acting on a gram for one second generates a velocity of a centimeter per second. The force of a dyne equals the weight of 1.02 m.g.

MEGADYNE equals 1,000,000 dynes.

ERG, the unit of *work* in the centimeter-gram-second (C. G. S.) system of measurement of force. It is the amount of work done by the unit of force, one dyne, acting through the unit of distance, one centimeter, in the unit of time, one second.

MEGAERG equals 1,000,000 ergs.

FOOT-POUND is the amount of work required to raise one pound one foot against the force of gravity in one second. One foot-pound is equal to 13.56 megaergs.

HORSE POWER is the force which is equal to 550 foot-pounds. One horse power is equal to 7458 megaergs.

ELECTRICAL UNITS.

OHM is the unit of resistance: that offered by a column of mercury at 0° C. 106.3 cm. long, weighing 14.4521 grams, and having a uniform cross-section.

AMPERE is the unit of current. An ampère is such a current as will deposit 0.001118 gram of silver per second from a neutral solution of silver nitrate.

VOLT is the unit of electromotive force. A volt is the force required to produce a current of one ampère after overcoming the resistance of one ohm.

COULOMB is the unit of quantity of electricity transferred by a current of one ampère in one second.

FARAD is the unit of capacity. It is the capacity of a condenser, charged to one volt by one coulomb.

JOULE is the unit of work. It is the amount of energy required to send one ampère of current through one ohm resistance in one second. One joule equals 10,000,000 dynes.

WATT is the unit of power. A watt is equivalent to the work done at the rate of one joule per second. One watt is 10,000,000 ergs or $\frac{1}{746}$ h.p.

HENRY is the unit of induction. It is the induction in a circuit when the electromotive force induced in this circuit is one volt, while the inducing current varies one ampère per second.

TABLE OF ELEMENTS.

Name.	Symbol.	Atomic Weight.	Name.	Symbol.	Atomic Weight.
Aluminum	Al	27.04	Neodymium	Nd	142.6
Antimony	Sb	119.6	Nickel	Ni	58.6
Argon	A	40.0 ?	Nitrogen	N	14.01
Arsenic	As	74.9	Osmium	Os	190.3
Barium	Ba	136.9	Oxygen	O	15.96
Beryllium[1]	Be	9.03	Palladium	Pd	106.35
Bismuth	Bi	208.9	Phosphorus	P	30.96
Boron	B	10.9	Platinum	Pt	194.3
Bromine	Br	79.76	Potassium	K	39.03
Cadmium	Cd	111.5	Praseodymium	Pr	139.4
Cæsium	Cs	132.7	Rhodium	Rh	102.9
Calcium	Ca	39.91	Rubidium	Rb	85.2
Carbon	C	11.97	Ruthenium	Ru	101.4
Cerium	Ce	139.9	Samarium	Sm	149.62
Chlorine	Cl	35.37	Scandium	Sc	43.97
Chromium	Cr	52.0	Selenium	Se	78.87
Cobalt	Co	58.6	Silicon	Si	28.3
Columbium[2]	Cb	93.7	Silver	Ag	107.66
Copper	Cu	63.18	Sodium	Na	23.0
Erbium	Er	166.0	Strontium	Sr	87.3
Fluorine	F	19.0	Sulphur	S	31.98
Gallium	Ga	69.9	Tantalum	Ta	182.0
Germanium	Ge	72.3	Tellurium	Te	125.0
Gold	Au	196.7	Terbium	Tb	159.7
Helium	He	4.0	Thallium	Tl	203.7
Hydrogen	H	1.0	Thorium	Th	231.9
Indium	In	113.6	Thulium	Tu	169.4
Iodine	I	126.53	Tin	Sn	118.8
Iridium	Ir	192.5	Titanium	Ti	48.0
Iron	Fe	55.88	Tungsten	W	183.6
Lanthanum	La	138.2	Uranium	U	238.8
Lead	Pb	206.4	Vanadium	V	51.1
Lithium	Li	7.01	Xenon	X	127.1
Magnesium	Mg	24.3	Ytterbium	Yb	172.6
Manganese	Mn	54.8	Yttrium	Yt	88.9
Mercury	Hg	199.8	Zinc	Zn	65.1
Molybdenum	Mo	95.9	Zirconium	Zr	90.4

[1] Also called Glucinum.
[2] Also called Niobium.

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NAME	DATE DUE